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CHEMISTRY 30

MODULE 5: CHEMISTRY IN BALANCE

Distance Learning

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Chemistry 30

Module 5

Chemistry in Balance

Distance Learning

Alberta EDUCATION

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Other	

Chemistry 30
Student Module
Module 5
Chemistry in Balance
Alberta Distance Learning Centre
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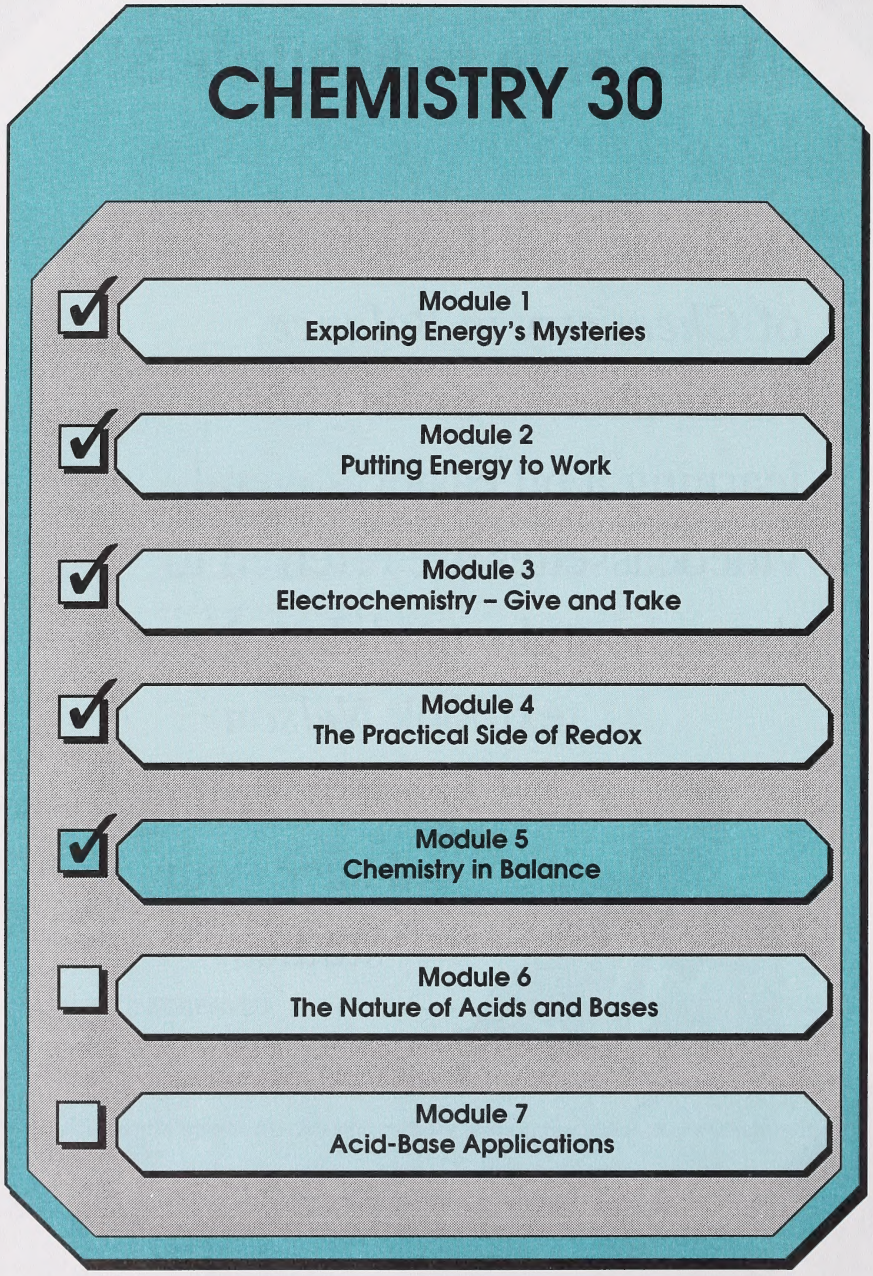
Welcome to Module 5!

We hope you'll enjoy your study of *Chemistry in Balance*. Remember, to make your learning a bit easier, several videocassettes are referred to throughout the modules. Your textbook *Nelson Chemistry* will also enhance your studies.

COURSE OVERVIEW

This course contains seven modules. The module you are working in is highlighted in deeper colour. It is recommended that you work through the modules in the order given since several concepts build on each other as you progress in the course.



MODULE 5

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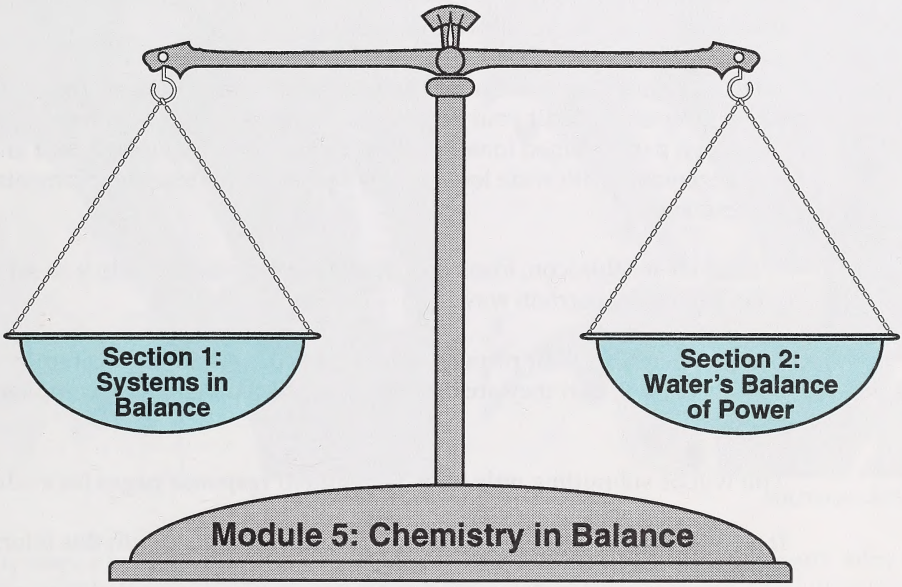
MODULE OVERVIEW

As a very young child playing on the teeter-totter, you found out quickly that you could not lift your older friend off the ground without help from some other friends. As your teeter-totter skills improved, you might have found that you could lift and balance your older friend by yourself by changing where you or your friend were sitting along the length of the teeter-totter. In this module, you will study another balancing act by considering chemical systems. You will discover how the balance in a chemical system can be affected in some way by changing conditions (like where you and your friend were sitting) or by adding or taking away parts (like extra friends). You will see how, in some cases, seemingly impossible tasks can be achieved.



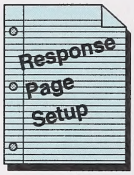
Evaluation

Your mark in this module will be determined by how well you complete the assignments at the end of each section. You must complete all assignments. In this module you are expected to complete three assignments. The mark distribution is as follows:

Section 1 Assignment	**30 marks**
Section 2 Assignment	**45 marks**
Final Module Assignment	**25 marks**
TOTAL	**100 marks**

When doing your assignments, work slowly and carefully. If you are having difficulties, go back and review the appropriate section.

Read all parts of your assignment carefully. Plan and do your rough work on your own paper. Revise and edit your responses; then set up your final copy for submission on your own paper. Lined looseleaf is recommended. Make sure your answers are neat and organized, with wide left margins and space for teacher comments after each assignment.



When you see this icon, ideas and details are provided to help you set up and organize your answer in a certain way.

Before submitting your responses, be sure to proofread them carefully to ensure they say what you want, that they are neat and clear, and that they are complete and missing no material.

You will be submitting **only** your **assignment response pages** for evaluation.

It is important to number and clearly identify each page with this information placed at the top.

Chemistry 30 – Module 5 Section # Assignment Page # Name and ID #

SECTION 1

Systems in Balance

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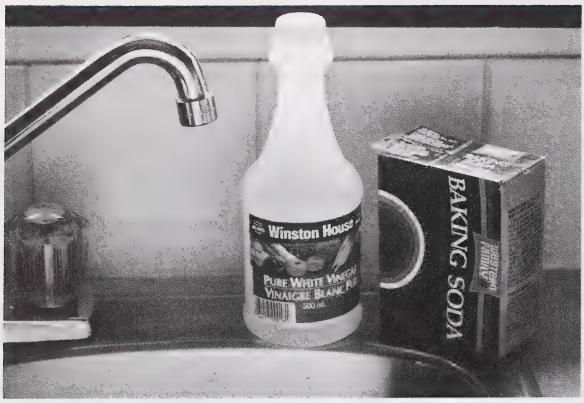
Why does a bottle of soft drink immediately start fizzing upon being opened? Why does an opened bottle of soft drink go *flat* faster when left in the sun than when left in the refrigerator? The answers to these questions may also reveal why a shaken bottle of pop produces a much more satisfying stream of liquid when pointed at an unsuspecting target. In chemical terms, opening the bottle disturbed a system in balance; the release of a gas in the form of bubbling is this same system trying to find a new balance. The balance in a system can be upset by several changes in conditions. By controlling these conditions, you can control the balance of the system.

In this section, you will be introduced to the term *equilibrium* in dealing with some typical chemical systems which are balanced to differing degrees. You will practise communicating the degree of balance in several different ways. Through investigations you will discover ways to upset systems in balance by altering the conditions in which the systems are found. You will also see some important industrial applications of equilibrium concepts.

ACTIVITY 1

Systems and Equilibrium



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Have you ever used vinegar and baking soda to clean the drains of your kitchen sink? The reaction is vigorous enough to scour the insides of the pipes, and is environmentally friendly as well. In the following investigation, using vinegar and baking soda, you will make a chemical system which you will manipulate in order to study how a system's balance can be affected by the presence of another substance.

Important Note – Disposal Information

Any chemicals used in investigations in this module should not be poured down the sink unless stated otherwise. Keep waste materials in closed, labelled containers. Contact your nearest high school or fire department for disposal information.

Investigation: Baking Soda and Vinegar

Science Skills

- ☐ A. Initiating
- ☑ B. Collecting
- ☑ C. Organizing
- ☑ D. Analysing
- ☐ E. Synthesizing
- ☐ F. Evaluating

Carefully follow the instructions given in this investigation. Pay careful attention to the required components, safety aspects, and applied science skills.

Problem

How can the balance of a chemical system be changed by the addition of another substance?

Materials

- baking soda
- vinegar
- bromothymol blue
- (2) 100 mL beakers
- metric measuring spoon
- 100 mL graduated cylinder
- thermometer
- eyedropper

Procedure

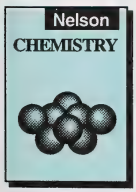
- Make a solution of baking soda in water using 2 mL (1/2 tsp.) of baking soda in 50 mL water.
- Stir the mixture until the baking soda is completely dissolved.
- Record the colour and appearance of the solution.
- Add 4 or 5 drops of bromothymol blue to the solution you formed.
- Record the colour and appearance of the new solution.
- Measure and record the temperature of your solution.
- Add 10 mL (2 tsp.) of vinegar to the solution you have made; stir the mixture.
- Record the colour and temperature of the resulting mixture and any other relevant observations.
- Continue adding 10-mL increments of vinegar to the sodium bicarbonate solution; record your observations until no further reaction occurs.

1. Construct an evidence table, perform the procedure, and record your observations.

Check your answers by turning to the Appendix, Section 1: Activity 1.

equilibrium – the condition of a closed system in which all observable properties remain constant at a constant temperature

The preceding investigation illustrates one example of a system at **equilibrium**. You upset the equilibrium by changing the conditions and you observed the results.



Equilibrium is not a new concept to you. In Chemistry 20, you studied solution equilibria. Before you answer the next questions that in part analyse the results of the preceding investigation, read pages 426 and 427 of your textbook to review some equilibrium concepts and to see some applications.

2. a. In the preceding baking soda/vinegar investigation, what was the colour of the sodium bicarbonate solution after you added bromothymol blue?

   b. Based on what you learned in Chemistry 20, this colour should indicate the presence of a _____.

macroscopic – that which is observable

3. Was the system at equilibrium before adding the vinegar? List three **macroscopic** properties that would confirm your answer.

4. List four empirical properties that can be measured for any system at equilibrium.

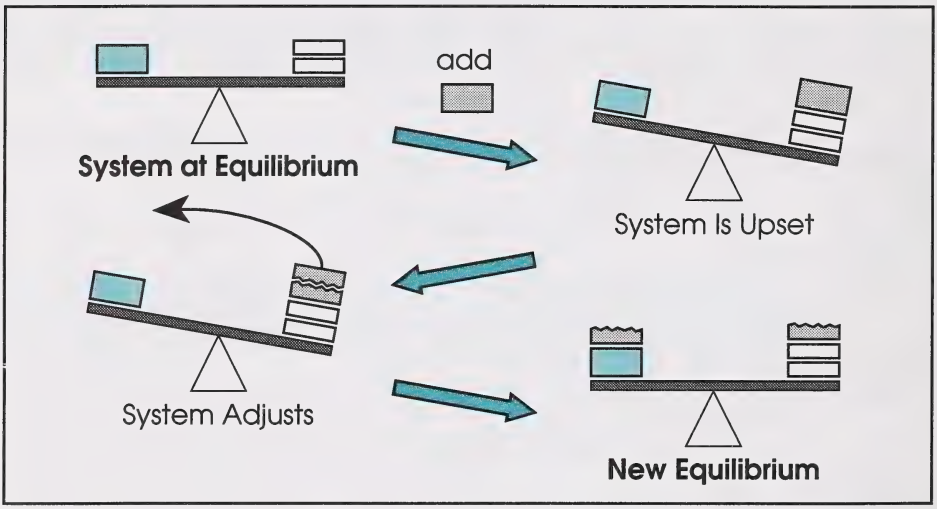
5. While adding vinegar, was the system still at equilibrium? Give three observations that support your answer.

6. A student performed the same investigation and tested the gas produced. The student found that the gas created a milky precipitate in the limewater diagnostic test. What gas was produced? How is this similar to opening a bottle of soft drink?

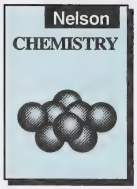
7. Why does the gas escape immediately in your investigation, but only after a soft drink bottle is opened?

8. In your investigation, what observations tell you that the reaction was finished?

Check your answers by turning to the Appendix, Section 1: Activity 1.



An equilibrium system, like other chemical systems, can be explained using the collision-reaction theory introduced in Chemistry 20. Chemical reactions depend on collisions between reactants resulting in the breaking and forming of chemical bonds.



Read page 174 of *Nelson Chemistry* to remind yourself of the collision-reaction theory and how it relates to stoichiometric calculations.

9. What four assumptions are made in the collision-reaction theory?

10. Which assumption of the theory was judged to be unacceptable in chapter 12 of your textbook (when you dealt with redox reactions)?

The collision-reaction theory assumes that all chemical reactions go to completion (that is, they are quantitative). This assumption of the theory has been proved wrong once, so is this assumption correct or not? You will test this assumption further in the investigation that follows.

Investigation 14.1: Extent of a Chemical Reaction

Nelson CHEMISTRY

Carefully follow the directions given for Investigation 14.1 on pages 427-428 of your textbook. Pay special attention to the required components, safety aspects, and applied science skills.

Guidelines

Science Skills

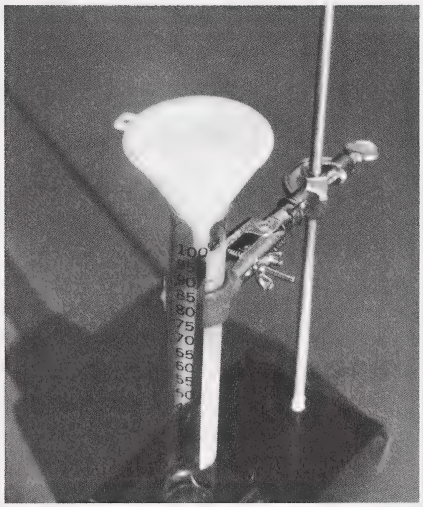
- ☑ A. Initiating
- ☑ B. Collecting
- ☑ C. Organizing
- ☑ D. Analysing
- ☐ E. Synthesizing
- ☑ F. Evaluating

- The solutions you will be using will all be 0.10 mol/L rather than 0.50 mol/L, 1.0 mol/L, or saturated.

Caution

- Your filtration apparatus can be set up similar to the one on the right, or similar to the one as shown on page 534 of *Nelson Chemistry*.

  If you feel the need, review the filtration procedure described on pages 534-535 of your textbook.



11. Outline your procedure in point form. (**Hint:** You will need to include the volumes of each solution used per trial.)

Check your answers by turning to the Appendix, Section 1: Activity 1.

12. Construct an evidence table, perform the procedure, and record your evidence.

13. Write the balanced equation, including states, for the reaction that occurs in your investigation. Which product was the precipitate?

14. For each of your trials, determine which reactant should have been completely used up (the limiting reagent) and which reactant should have been in excess.

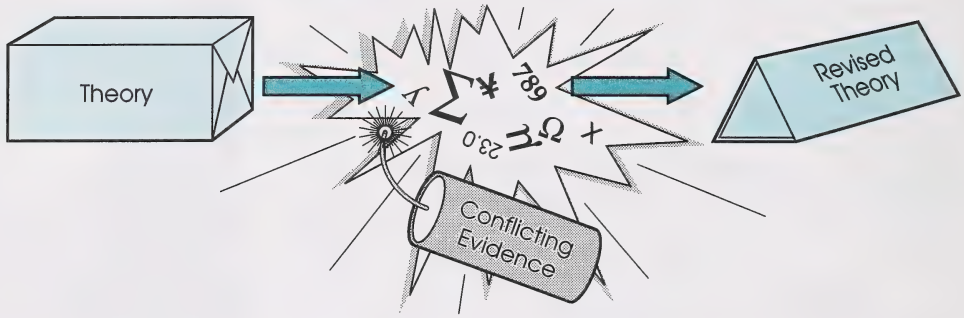
15. What do the results of the diagnostic tests indicate?

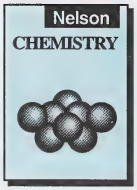
16. Evaluate your results. Did your analysis confirm that reactions always react quantitatively?

17. If you were unable to check your evidence for reliability with other students, what could you do instead?

Check your answers by turning to the Appendix, Section 1: Activity 1.

Is the collision-reaction theory valid?





When scientists gather new evidence that does not agree with accepted laws, they have to consider revising the old laws. Contradictory experimental evidence such as you just discovered has led to a new application of the collision-reaction theory in which equilibrium systems are a competition between reactants forming products and products forming reactants. Read the top half of page 429 in your textbook to obtain an understanding of the competition that occurs between reactants and products; then answer the following questions.

18. Use the appropriate words or terms to fill in the blanks in the following paragraphs. (Some words or terms may be used more than once; some not at all.)

- $I_{2(aq)}$
- dynamic
- closed
- constant
- vapour
- $I_{2(s)}$
- reverse
- solubility
- balance
- rate
- equilibrium
- react
- phase
- liquid
- chemical

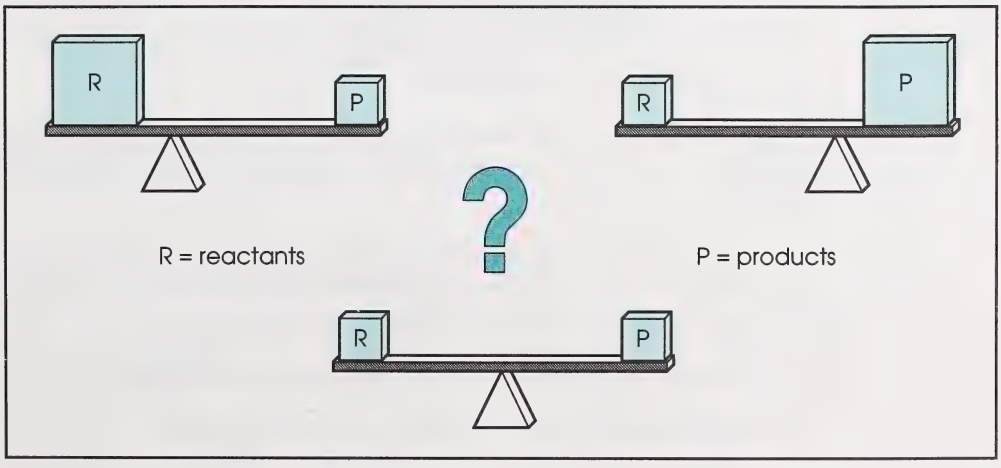
The presence of reactants in a system that was assumed to react quantitatively to produce products is evidence that the ____a.____ reaction has occurred to some extent. This ____b.____ between the forward and reverse rates of reaction is known as ____c.____. This balance can only exist in a ____d.____ environment to ensure that neither products nor reactants escape.

Systems can have different types of equilibria, including chemical, ____e.____, and ____f.____. In a saturated solution of iodine, for example, the rate of $I_{2(s)}$ forming ____g.____ is equal to the ____h.____ of $I_{2(aq)}$ forming ____i.____. This is an example of a ____j.____ equilibrium. In a sealed glass of water, ____k.____ water evaporates at the same ____l.____ as water ____m.____ condenses. Sodium sulphate and calcium chloride solutions reacting to form products with calcium sulphate solid and sodium chloride solution reacting to re-form the original reactants is an example of a ____n.____ equilibrium.

19. In a reaction system, if it is easier to carry out the forward reaction than the reverse reaction (that is, factors favour the forward reaction), would you expect equal amounts of products and reactants at equilibrium? Explain.

Check your answers by turning to the Appendix, Section 1: Activity 1.

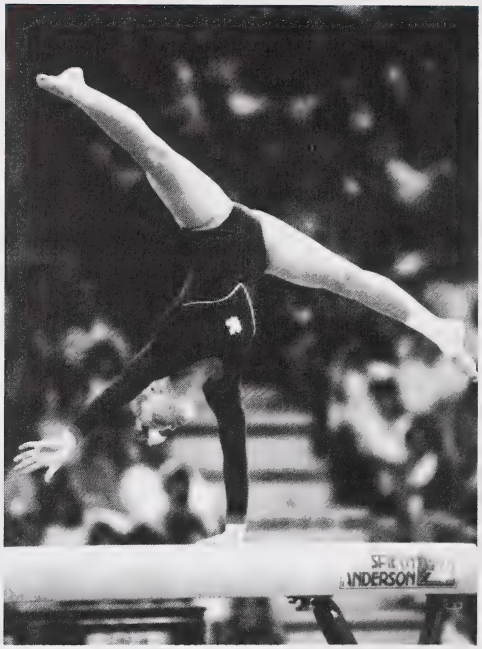
You have discovered that in some reactions there appears to be a balance between reactants and products. The next activity will help you to discover where this balance lies. Are there equal amounts of reactants and products in an equilibrium, more products than reactants, or more reactants than products? Are there any other possibilities?



ACTIVITY 2

What Is a Reaction in Balance?

Keeping your balance in gymnastics can often be quite tricky. Do you find it's the same way with balancing your chequing account? Most people use a simple formula of subtracting what was spent from what was put in. But a balanced chequing account means different things to different people. Money goes in, money comes out, and at the end of the month, some people have a large negative balance, some a large positive balance, and most will have a balance close to zero.



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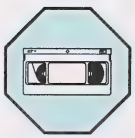
Chemical systems are much the same. Reactants go in, products come out; some reactions produce very little product, some reactions go to completion (use up all the reactants), and most end up somewhere in the middle. In this activity you will learn how to calculate the degree of balance in a chemical system and how to communicate this balance to others, much the same way as your bank communicates your account balance by sending you a monthly bank statement.



To learn about chemical systems with different balances, watch the video program called Dynamic Equilibrium from the video series *Chemical Equilibrium*. Familiarize yourself with the following questions prior to watching the program. This will help you to focus on the main ideas while you are viewing. You may periodically have to stop the tape in order to record your answers.

1. Explain how the colour changes were created in the dinitrogen tetraoxide-nitrogen dioxide mixture. What happened when both containers were left at the same temperature?

2. a. When hydrogen gas and iodine gas are mixed and then heated, they react to form _____. Evidence to indicate this reaction occurs is the _____ of the purple colour associated with _____.

   b. What does the pink colour left in the reaction mixture indicate?

   c. What does the constancy of the pink colour suggest?

3. a. According to the kinetic molecular theory, what three types of motion are molecules assumed to have?

   b. In order for two molecules to react, what two conditions must they meet?

4. a. Briefly describe how the hydrogen iodide equilibrium develops.

   b. At what point, considering the forward and reverse reactions, is equilibrium reached?

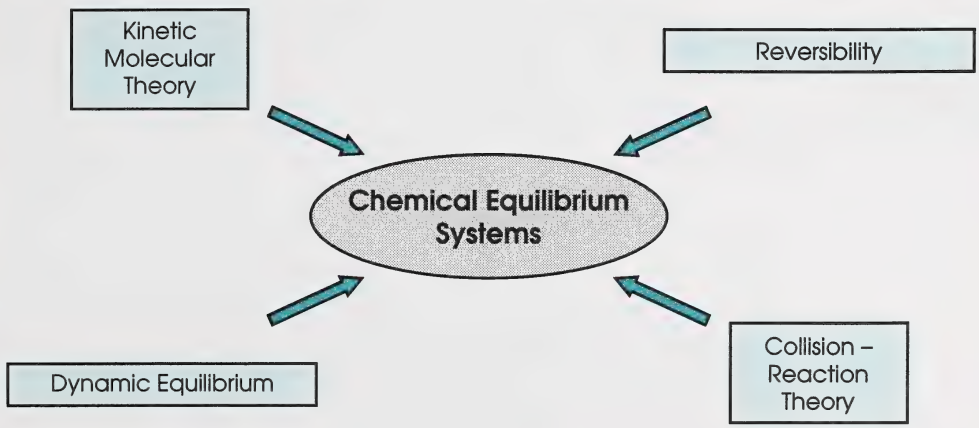
   c. What kind of equilibrium is this?

5. Reproduce in your notebook the graph of quantity (concentration) versus time for hydrogen reacting with iodine. Indicate on the graph where equilibrium conditions exist.

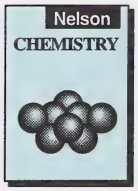
6. How is a dynamic equilibrium illustrated in a reaction equation?

7. From the video briefly describe the iodine experiment that indicated a chemical equilibrium system is dynamic.

Check your answers by turning to the Appendix, Section 1: Activity 2.



Now that you have seen that not all chemical reactions go to completion, that chemical equilibrium exists, and that the balance between reactants and products is not the same for all systems, you need to find a way to communicate where the balance lies. In the preceding video, you discovered that the reaction of hydrogen and iodine gases to produce hydrogen iodide has easily observable properties. Your textbook also uses this reaction to examine some quantitative properties of a system at equilibrium.



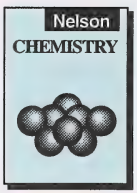
Read the bottom of page 429 and the top half of page 430 in your textbook. Table 14.1 gives initial and equilibrium concentrations for three hydrogen-iodine equilibria. At first glance, there appears to be no pattern in the data. Can you see a pattern?

8. a. Describe the initial system conditions in general terms for each of the three systems given in Table 14.1.

   b. Describe the final conditions (equilibrium system concentration) in general terms for each of the three systems.

Check your answers by turning to the Appendix, Section 1: Activity 2.

An important point for you to realize is that all three systems shown in Table 14.1 reached an equilibrium state in spite of starting from different points. How are these three different equilibrium states related? Is there some mathematical relationship that exists to express the position of the equilibrium? A commonly used measure of comparison is percentage.



You already know how test marks are converted to a percentage from a raw score to give a common basis for comparison. In much the same manner that you might describe how you do on a test by converting your mark to a percentage, chemists describe the extent of a reaction by calculating the **percent reaction** or percent yield at equilibrium. Percent reaction at equilibrium is the chemical equivalent of a percentage grade. How is it calculated? Find out by reading the section Percent Reaction at Chemical Equilibrium on page 430 in your textbook. Then complete the following questions.

percent reaction – the equilibrium concentration of product in an equilibrium as compared to the maximum possible concentration, expressed as a percentage

9. Fill in the missing information in the following comparison.

$$\text{test percentage} = \frac{\text{your mark}}{\text{total marks available}} \times 100\%$$

$$\text{percent reaction} = \frac{\qquad}{\qquad} \times 100\%$$

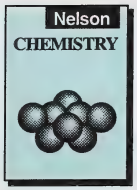


Table 14.2 on page 430 of your textbook gives the values necessary to calculate the percent reactions for each of the hydrogen-iodine systems mentioned earlier. The following sample calculation shows you how those numbers were calculated.

Example

System	Equilibrium [HI] (mmol/L)	Maximum Possible [HI] (mmol/L)	Percent Reaction (%)
2	2.10	2.70	77.8

Equilibrium [HI]: This value, 2.10 mmol/L, came from the experiment results given in Table 14.1 in your textbook. The value was obtained by measurement when System 2 reached equilibrium.

Maximum Possible [HI]: This value, 2.70 mmol/L, is obtained by using the balanced chemical equation for the system and the initial $\left[HI_{(g)}\right]$.

$$H_{2(g)} + I_{2(g)} \rightleftarrows 2\,HI_{(g)}$$

The mole to mole ratio from the equation indicates that 2 mol of $HI_{(g)}$ will be produced for every 1 mol of either reactant that reacts. Since System 2 had 0.50 mmol/L each of $H_{2(g)}$ and $I_{2(g)}$, the maximum $\left[HI_{(g)}\right]$ that could be produced is $2 \times 0.50 \text{ mmol/L} = 1.00 \text{ mmol/L}$. This value would be added to the $\left[HI_{(g)}\right]$ initially present (1.70 mmol/L) to give the total maximum $\left[HI_{(g)}\right]$.

$$1.00 \text{ mmol/L} + 1.70 \text{ mmol/L} = 2.70 \text{ mmol/L}$$

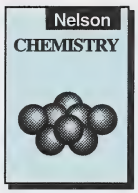
Percent Reaction: This value, 77.8%, is calculated using the percent reaction formula.

$$\begin{aligned}\text{percent reaction (percent yield)} &= \frac{\text{equilibrium } \left[HI_{(g)}\right]}{\text{maximum } \left[HI_{(g)}\right]} \times 100\% \\ &= \frac{2.10 \text{ mmol/L}}{2.70 \text{ mmol/L}} \times 100\% \\ &= 77.8\%\end{aligned}$$

10. Calculate the percent reaction for each of the hydrogen-iodide systems in the following table (all concentrations are in mmol/L). The first one is done for you.

THE HYDROGEN-IODIDE SYSTEM AT 448°C

System	Initial $[H_{2(g)}]$	Initial $[I_{2(g)}]$	Initial $[HI_{(g)}]$	Maximum $[HI_{(g)}]$	Equilibrium $[HI_{(g)}]$	Percent Reaction (%)
a	2.0	2.0	1.0	**5.0**	3.9	**78**
b	10.0	10.0	0.0		15.6	
c	7.50	7.50	2.00		13.3	
d	0.40	0.40	1.60		1.87	
e	0.00	0.00	1.50		1.18	
f	1.20	1.20	1.00		2.65	



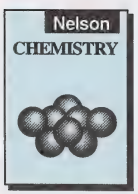
Being able to calculate the percent reaction is just the first step in communicating the balance of a chemical reaction. Read page 431 in your textbook to learn how to combine the percent reaction of an equilibrium system with the chemical equation of the system for communicating the balance or position of a chemical equilibrium.

$$\text{reactants} \overset{\%}{\rightleftarrows} \text{products} \qquad t = \_\_\_\,°C$$

When you are writing a chemical equation for an equilibrium system using percent reaction, be sure to include

- the double arrow, $\rightleftarrows$, to indicate an equilibrium
- the percent reaction over the arrow to illustrate the extent or position of the equilibrium
- the temperature for which this percent reaction is valid (otherwise assume SATP)

Practise communicating equilibrium positions by doing the following question.



11. Do textbook questions 7 (a-d) and 8 on pages 431-432 of *Nelson Chemistry*.

Expressing the extent of a chemical equilibrium using percent reaction is a quantitative description. Table 14.3 on page 431 of your textbook includes the qualitative descriptors that are also used to describe equilibrium positions. Use Table 14.3 to answer the following question.

12. For each of the reactions described in textbook question 7 (a-d) that you just answered, classify the reaction as quantitative, products favoured, or reactants favoured.

Check your answers by turning to the Appendix, Section 1: Activity 2.

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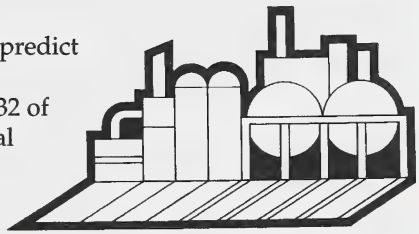
Although percent reaction is a valid way to express equilibrium position, it does not allow a chemist to predict the equilibrium position given a new set of initial conditions. In the following lab exercise, on page 432 of your textbook, you will develop a new mathematical relationship that will serve this function for a particular chemical system.



Science Skills

- ☐ A. Initiating
- ☐ B. Collecting
- ☐ C. Organizing
- ☑ D. Analysing
- ☐ E. Synthesizing
- ☐ F. Evaluating

Lab Exercise 14A: The Synthesis of an Equilibrium Law

13. Read Lab Exercise 14A and complete the analysis of the investigation report. Pay special attention to the applied science skill. Construct a table similar to the following. Calculate the four relationships shown for each trial as part of your analysis.

equilibrium law – an expression equal to the product of the equilibrium concentrations of the chemical products divided by the product of the equilibrium concentrations of the chemical reactants, where each concentration is raised to the power of the coefficient of that entity

THE IRON(III)-THIOCYANATE REACTION SYSTEM

Calculation	Trial 1	Trial 2	Trial 3	Trial 4	Trial 5
$[Fe^{3+}][SCN^-][FeSCN^{2+}]$					
$[Fe^{3+}] + [SCN^-] + [FeSCN^{2+}]$					
$\frac{[Fe^{3+}][SCN^-]}{[FeSCN^{2+}]}$					
$\frac{[FeSCN^{2+}]}{[Fe^{3+}][SCN^-]}$					

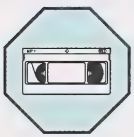
14. Based on your answer to the previous question, which, if any, relationships give a near constant value (within experimental error)?

Check your answers by turning to the Appendix, Section 1: Activity 2.

The two relationships that gave you constant values are related mathematically. They are reciprocals of one another. By convention (to make sure that there is no confusion) chemists have decided to always place product concentrations over reactant concentrations. This mathematical relationship that you found in Lab Exercise 14A is known as the equilibrium law; the constant numerical value is called the **equilibrium constant, *K***.

equilibrium constant (*K*) – the value obtained from the mathematical combination of equilibrium concentrations using the equilibrium law

$$K_{\text{equilibrium constant}} \rightarrow \frac{\text{products}}{\text{reactants}}$$



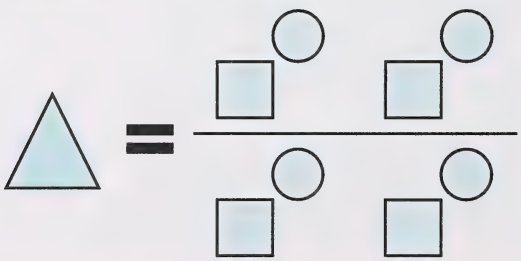
To see how similar laws are developed for other equilibrium systems, watch the program called The Equilibrium Constant from the video series *Chemical Equilibrium*. This program does an excellent job of explaining the equilibrium constant, the conventions used, how to write the equilibrium law, and what the size of the constant tells you about the equilibrium. Watch the program and then answer the following questions.

15. The amounts of reactants and products in equilibrium law expressions are measured in _____.

16. When writing the equilibrium law, the _____ are always in the numerator and the _____ are always in the denominator.

17. The hypothetical equation $w\text{Q} + x\text{R} \rightleftarrows y\text{S} + z\text{T}$ represents a chemical system at equilibrium.

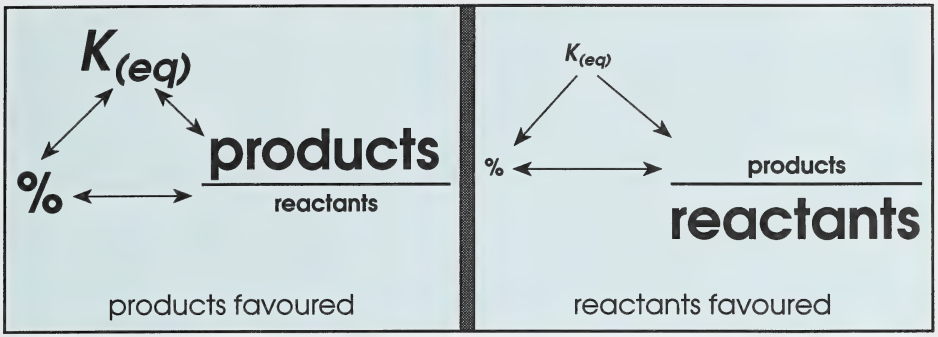
    Place the letters w, x, y, z, Q, R, S T, and K in the appropriate places in a graphic organizer similar to the one illustrated.



Check your answers by turning to the Appendix, Section 1: Activity 2.

The qualitative descriptors you used with percent reaction can also be used with the equilibrium constant.

18. The following graphic shows the relationships between percent reaction, K, and relative amounts of products and reactants.



A large K value means a _____ percent reaction and that _____ are favoured. A small K value means a _____ percent reaction and that _____ are favoured.

19. Complete the reaction classification column of the following table using the terms *quantitative*, *products favoured*, *reactants favoured*, or *balanced*.

K Value	Classification
5.6	
0.96	
1.0×10^{12}	
3.1×10^{-8}	

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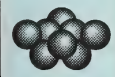


Like all mathematical formulas, the equilibrium law can be used to find any variable if all others are known. You can find K if you know all the equilibrium concentrations. If you know K, you can find one missing concentration. You could, for example, use the equilibrium data from Table 14.1 on page 430 of your textbook to calculate the K value for the hydrogen-iodine system; then use the calculated value to find a concentration.

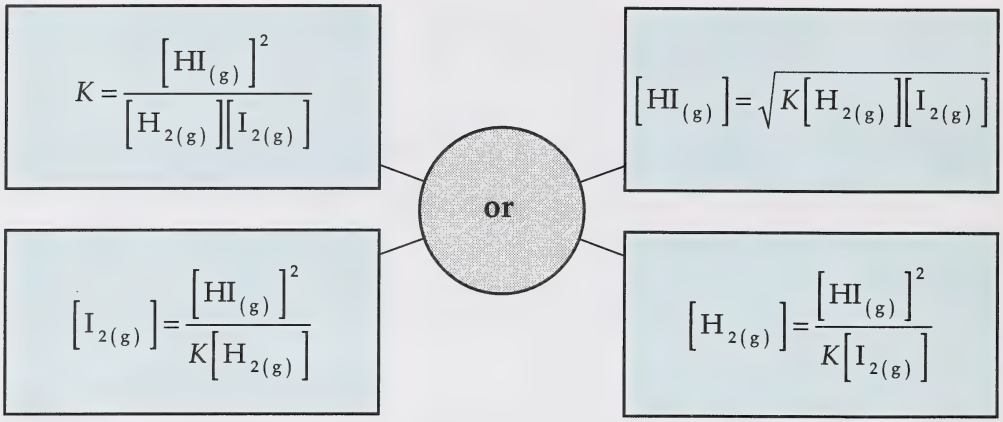
20. Write the equilibrium law for the hydrogen-iodine system, as determined from the values from textbook Table 14.1 (page 430).

21. a. Substitute the equilibrium values for each of the systems from Table 14.1 into the equilibrium law (expression). Calculate K for each system at 448°C.

    b. Calculate the average K for the systems at 448°C.

Check your answers by turning to the Appendix, Section 1: Activity 2.

The equilibrium law can be used to calculate any of the equilibrium concentrations if the K value and other concentrations are known.



22. Use the K value you calculated in question 21.b. to complete the following table of equilibrium conditions for the hydrogen-iodine system at 448°C. The first two are done as examples.

EQUILIBRIUM CONDITIONS FOR HYDROGEN-IODINE SYSTEM AT 448°C

System	$[H_{2(g)}]$ (mol/L)	$[I_{2(g)}]$ (mol/L)	$[HI_{(g)}]$ (mol/L)
a.	0.40	**0.050**	1.00
b.	0.16	9.00	**8.5**
c.	———	0.333	3.750
d.	3.56	———	22.4
e.	4.0x	9.0x	———

The calculations for Systems a. and b. are as follows:

a. $$[I_{2(g)}] = \frac{[HI_{(g)}]^2}{K[H_{2(g)}]}$$
$$= \frac{(1.00\ \text{mol/L})^2}{50 \times 0.40\ \text{mol/L}}$$
$$= 0.050\ \text{mol/L}$$

b. $$[HI_{(g)}] = \sqrt{K[H_{2(g)}][I_{2(g)}]}$$
$$= \sqrt{50 \times 0.16\ \text{mol/L} \times 9.00\ \text{mol/L}}$$
$$= 8.5\ \text{mol/L}$$

You have found that at 448°C the average K value for the hydrogen-iodide system, $H_{2(g)} + I_{2(g)} \rightleftarrows 2\,HI_{(g)}$, is 50, and that if two of the equilibrium concentrations are known, then the other can be found.

Another equilibrium situation that may arise is being asked to find the value of K when you are given the initial concentrations of reactants and one equilibrium concentration.

Example

Question:

The following equilibrium is in a 2.00 L sealed flask at a constant temperature.

$$H_{2(g)} + I_{2(g)} \rightleftarrows 2\,HI_{(g)}$$

Initially, 0.200 mol of $H_{2(g)}$ and 0.200 mol of $I_{2(g)}$ were introduced into the flask. When an equilibrium was reached, it was found that the equilibrium concentration of $I_{2(g)}$ was 0.020 mol/L. Calculate the value of K for this equilibrium.

Answer:

Step 1 Write the equilibrium law expression.

$$K = \frac{[HI]^2}{[H_2][I_2]}$$

In order to solve for K, you must first find the *equilibrium* concentrations of all participants. This will take a couple of steps.

Step 2 Calculate the initial concentrations of $H_{2(g)}$ and $I_{2(g)}$. (There is no product, initially.)

$$\text{initial } [H_{2(g)}] = \frac{0.200 \text{ mol}}{2.00 \text{ L}} = 0.100 \text{ mol/L}$$

← number of moles given
← 2.00 L flask

$$\text{initial } [I_{2(g)}] = \frac{0.200 \text{ mol/L}}{2.00 \text{ L}} = 0.100 \text{ mol/L}$$

← number of moles given
← 2.00 L flask

Step 3

Find the equilibrium concentrations by comparing the initial concentrations to the amount each participant has changed. This information can be organized in a table.

	$H_{2(g)}$	+	$I_{2(g)}$	$\rightleftarrows$	$2\ HI_{(g)}$
initial concentrations (mol/L)	0.100		0.100		0.000
changed concentrations (mol/L)	– 0.080		– 0.080		2(0.080)
equilibrium concentrations (mol/L)	0.020		0.020		0.160

[ ] – from given information

How were the entries in the table obtained?

Initial concentrations – You calculated these in Step 2.

Changed concentrations – These values reflect how the concentration of each substance has changed to come to an equilibrium. Since the equilibrium $\left[I_{2(g)}\right]$ was given (0.020 mol/L), you can calculate the concentration of $I_{2(g)}$ that reacted or changed.

$$\text{initial } \left[I_{2(g)}\right] - \text{equilibrium } \left[I_{2(g)}\right] = \text{reacted } \left[I_{2(g)}\right]$$
$$0.100 \text{ mol/L} - 0.020 \text{ mol/L} = 0.080 \text{ mol/L}$$

The changed concentration of all participants are related by their mole ratios from the chemical equation: $1\ H_{2(g)} : 1\ I_{2(g)} : 2\ HI_{(g)}$. Therefore, since the changed $\left[I_{2(g)}\right]$ is 0.080 mol/L, the changed $\left[H_{2(g)}\right]$ must also be 0.080 mol/L $\left(\mathbf{1}\ I_{2(g)} : \mathbf{1}\ H_{2(g)}\right)$, and the changed (produced) $\left[HI_{(g)}\right]$ must be 2×0.080 mol/L = 0.160 mol/L $\left(\mathbf{1}\ I_{2(g)} : \mathbf{2}\ HI_{(g)}\right)$.

Equilibrium concentrations – Because the $\left[H_{2(g)}\right]$ and $\left[I_{2(g)}\right]$ decreased by 0.080 mol/L, you would subtract the value from their initial concentrations. Because the $\left[HI_{(g)}\right]$ increased by 0.160 mol/L, this value is added to its initial concentration.

Step 4 Substitute the equilibrium values into the equilibrium law expression and solve for *K*.

$$K = \frac{[HI]^2}{[H_2][I_2]}$$

$$= \frac{(0.160)^2}{(0.020)(0.020)}$$

$$= 64$$

The equilibrium constant value is 64. (Remember, units are not usually attached to equilibrium constants.)

Important points to remember:

- **All** values you deal with in the table you constructed have concentration units – mol/L.
- The **changed concentrations** are related by their **mole ratios**. Also, for the changed concentrations, **all** the reactants decrease and **all** the products increase if the reaction or equilibrium is shifted to the right. If the shift is towards the left, the opposite occurs.
- Only **equilibrium concentrations** are entered in an equilibrium law expression.

Practise solving the various types of equilibrium problems by completing the following questions.

23. Hydrogen gas can be produced industrially according to the following equilibrium reaction at 500°C.

$$CO_{(g)} + H_2O_{(g)} \rightleftarrows CO_{2(g)} + H_{2(g)}$$

At equilibrium, the following concentrations are found:

- $\left[CO_{(g)}\right] = 0.296 \text{ mol/L}$
- $\left[CO_{2(g)}\right] = 0.100 \text{ mol/L}$
- $\left[H_2O_{(g)}\right] = 0.0249 \text{ mol/L}$
- $\left[H_{2(g)}\right] = 0.300 \text{ mol/L}$

Calculate the value for the equilibrium constant, *K*.

24. The equilibrium constant, *K*, for the following reaction is 0.50 at a certain temperature.

$$SO_{3(g)} + NO_{(g)} \rightleftarrows NO_{2(g)} + SO_{2(g)}$$

The equilibrium concentrations for three of the participants are as follows:

- $\left[SO_{3(g)}\right] = 0.48$ mol/L
- $\left[NO_{2(g)}\right] = 0.15$ mol/L
- $\left[SO_{2(g)}\right] = 0.40$ mol/L

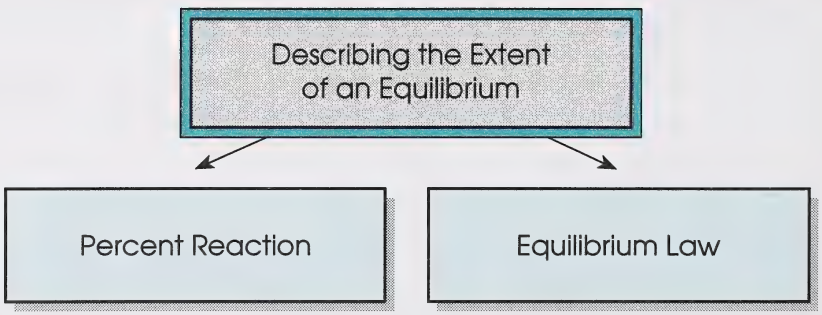
Calculate the equilibrium $\left[NO_{(g)}\right]$.

25. At a constant temperature, 0.750 mol of $HBr_{(g)}$ was placed in a 1.50 L container. The $HBr_{(g)}$ decomposed according to the following reaction:

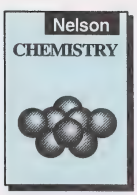
$$2\,HBr_{(g)} \rightleftarrows H_{2(g)} + Br_{2(g)}$$

At equilibrium, the concentration of $H_{2(g)}$ was found to be 0.130 mol/L. Calculate the value of the equilibrium constant *K*.

Check your answers by turning to the Appendix, Section 1: Activity 2.



You now have two mathematical means of communicating the extent of an equilibrium. Which way, if either, is better? Are there circumstances in which one method is better than the other?



Read pages 432-433 of your textbook to review the concepts covered so far (percent reaction and the equilibrium constant) and to discover some of the limitations in both these methods of communicating the extent of an equilibrium. Then answer the following questions.

26. Complete the following table by summarizing the advantages and disadvantages of each method of communicating a chemical equilibrium.

Method	Advantages	Disadvantages
Percent reaction		
Equilibrium constant		

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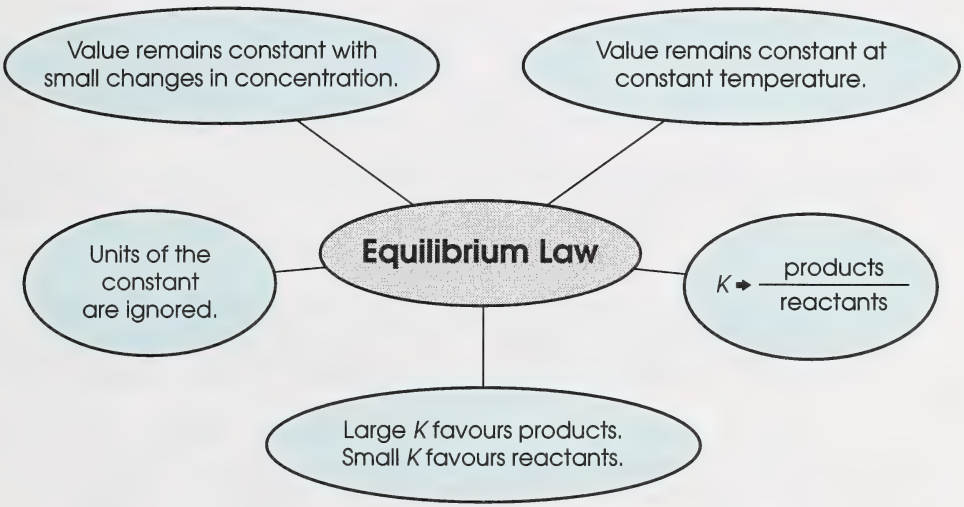
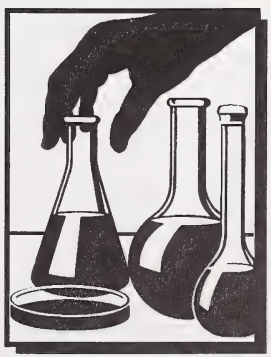
27. Do textbook questions 9 and 10 on page 434 of *Nelson Chemistry*.

Check your answers by turning to the Appendix, Section 1: Activity 2.


Value remains constant with small changes in concentration.
Value remains constant at constant temperature.
Units of the constant are ignored.
Equilibrium Law
$K \rightarrow \frac{\text{products}}{\text{reactants}}$
Large K favours products.
Small K favours reactants.

In this activity you have discovered ways of measuring and expressing the extent of a chemical reaction. In the next activity you will examine how changing the conditions of an equilibrium system can change the position of equilibrium.



ACTIVITY 3

Upsetting the Balance

What if the manufacturers of soft drinks were not able to get as much $CO_{2(g)}$ in their beverage containers as they now do? Left on its own, the amount of carbon dioxide that would naturally dissolve in water certainly would not provide the soda pop with enough fizz to convince consumers to drink a soft drink rather than ordinary flavoured water.

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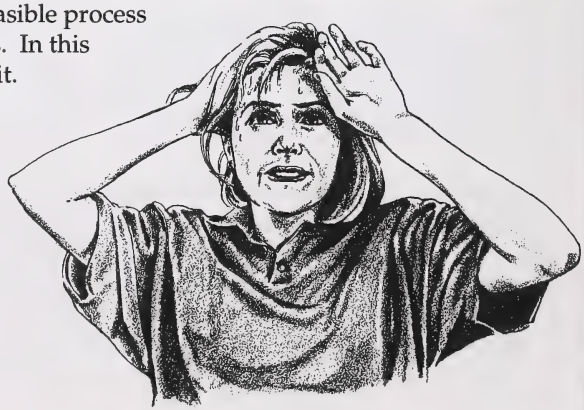
In Activity 2 you discovered that the equilibrium law and equilibrium constant give you an idea of the equilibrium position of a system.

1. How do the manufacturers of soft drinks manipulate the solution equilibrium so that more carbon dioxide will dissolve in the soft drink solution?

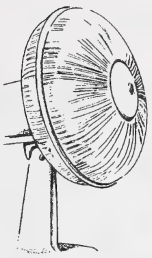
Check your answers by turning to the Appendix, Section 1: Activity 3.

There are other ways of manipulating equilibrium positions, and industrial chemists earn their living by technologically manipulating the equilibrium state of a chemical system to ensure an economically feasible process for producing their desired products. In this activity, you will learn how they do it.

How do you react to stress?



If it is too cold, for instance, you most likely go inside or put on warmer clothing. Too hot? You may turn on the air conditioning, open the windows, or go swimming. The point is that you do something to try and maintain your optimal equilibrium temperature. Chemical systems handle stresses in much the same way. The video series *Chemical Equilibrium* contains two programs, Reaction Tendencies and The Equilibrium Constant, that illustrate chemical systems under stress and how the systems overcome it.



Note: Even though you have already seen The Equilibrium Constant, it is good for you to view the program here to review the relationship between Le Châtelier's principle and the equilibrium law.

Familiarize yourself with the following questions prior to watching the programs to help you focus on the main ideas while you are viewing. You may periodically have to stop the tape in order to record your answers.

2. State Le Châtelier's principle.

3. Write the equation for the production of ammonia from hydrogen and nitrogen including a heat term.

4. The formation of ammonia from hydrogen and nitrogen is an _____ reaction; the decomposition of ammonia is an _____ reaction.

5. The following paragraph describes the effects of heat changes on the hydrogen-nitrogen-ammonia system. Choose the appropriate term from each set of brackets to fill in the blanks.

   When heat is added to a closed container of ammonia, nitrogen, and hydrogen at equilibrium, the equilibrium will shift to _____ (use up/produce) energy and relieve the stress. Since the _____ (production/decomposition) of ammonia uses up energy, the _____ (products/reactants) are favoured by the addition of heat.

6. Does the addition of heat always favour (shift towards) the reactants in all equilibrium reactions? Explain.

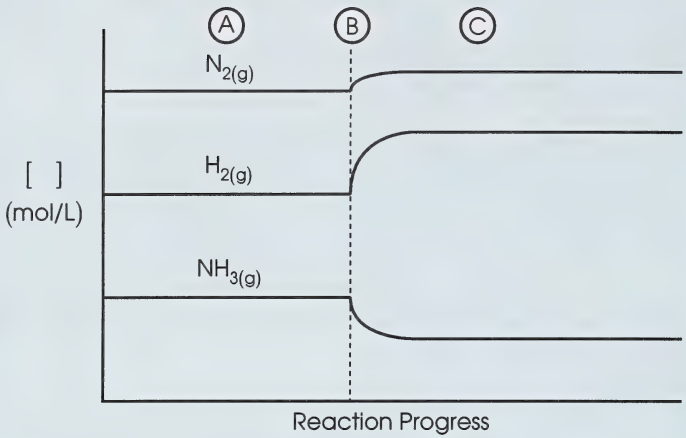
7. What property of the reaction determines which direction the reaction shifts (is favoured) with the addition of heat?

Check your answers by turning to the Appendix, Section 1: Activity 3.

The following example of an equilibrium graph is a visual method of representing the changes that occur when an equilibrium is upset.

Energy Change Effects on the Hydrogen-Nitrogen-Ammonia System

$$N_{2(g)} + 3\,H_{2(g)} \rightleftarrows 2\,NH_{3(g)} + 92\ kJ$$



Note: The graph is only intended to show qualitative changes (increase/decrease) rather than actual quantitative values (concentrations).

The relative starting points of each substance is chosen only for ease of illustration. What is important are the changes that occur.

(A) The system is at equilibrium. The forward and reverse reaction rates are equal. The concentration of each substance remains constant.

(B) Heat energy is added to the system.

(C) The system shifts in favour of the reverse reaction (endothermic direction) to try to use up the added heat energy. Therefore, the concentrations of $N_{2(g)}$ and $H_{2(g)}$ increase and the concentration of $NH_{3(g)}$ decreases. The system comes to a new equilibrium. The concentrations of all substances are again constant, but at different concentrations than the original equilibrium.

8. The following paragraph describes the effects of pressure changes on the hydrogen-nitrogen-ammonia system. Choose the appropriate term from each set of brackets to fill in the blanks.

   The _____ (production/decomposition) of ammonia results in a decrease in the total number of moles in a closed system and therefore a _____ (decrease/increase) in the total pressure of the system. When the pressure of an ammonia system is increased, the equilibrium will shift to _____ (decrease/increase) the total pressure and will cause the _____ (production/decomposition) of ammonia to relieve the stress.

9. Does an increase in pressure always force an equilibrium to produce more products? Explain.

10. What characteristic of the reaction determines whether the equilibrium will shift with an increase in pressure?

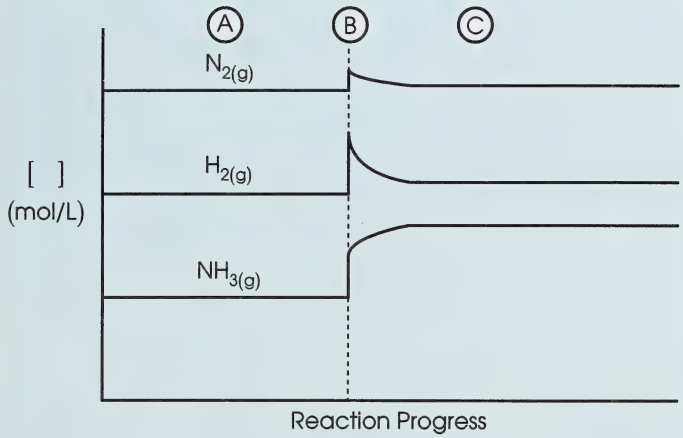
11. Why does an increase in pressure not shift an equilibrium system like $H_{2(g)} + I_{2(g)} \rightleftarrows 2\ HI_{(g)}$?

12. When studying the effects of increasing pressure on the ammonia system, why was it necessary to keep the system at a constant temperature?

Check your answers by turning to the Appendix, Section 1: Activity 3.

Pressure Change Effects on the Hydrogen-Nitrogen-Ammonia System

$$N_{2(g)} + 3\ H_{2(g)} \rightleftarrows 2\ NH_{3(g)} + 92\ kJ$$



Note: It is assumed that pressure changes are created by changing the size of the system's container. For example, an increase in pressure is achieved by decreasing the size of the container.

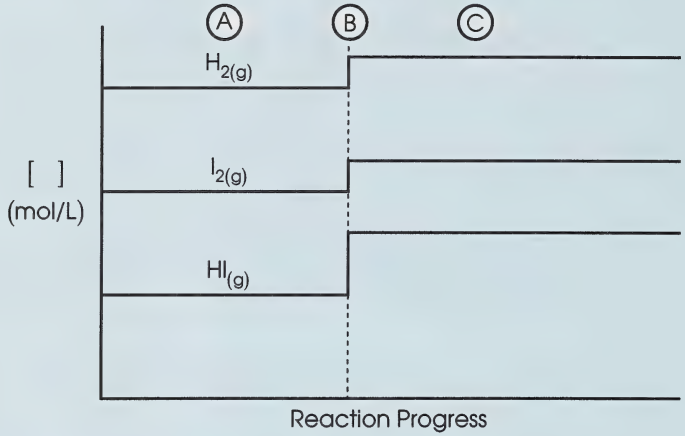
(A) The system is at equilibrium.

(B) The pressure of the system is increased by immediately decreasing the container volume. All concentrations immediately increase, proportional to their mole ratios.

(C) The system tries to relieve the increase in pressure by making fewer molecules (pressure is directly related to number of molecules). This means the system shifts towards the products (reactant side – 4 molecules; product side – 2 molecules). The system then comes to a new equilibrium.

Pressure Change Effects on the Hydrogen-Iodine-Hydrogen Iodide System

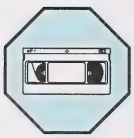
$$53\ kJ + H_{2(g)} + I_{2(g)} \rightleftarrows 2\ HI_{(g)}$$



(A) The system is at equilibrium.

(B) The pressure on the system is increased by immediately decreasing the container size. The concentrations of all participants immediately increase, proportional to their mole ratios.

(C) There is no shift in equilibrium because the system cannot shift to relieve the pressure by creating fewer molecules (reactant side – 2 molecules; product side – 2 molecules). The system is at a new equilibrium.



The following question relates to the effects of changes in concentration of substances in an equilibrium system, as discussed in the first part of the video program The Equilibrium Constant.

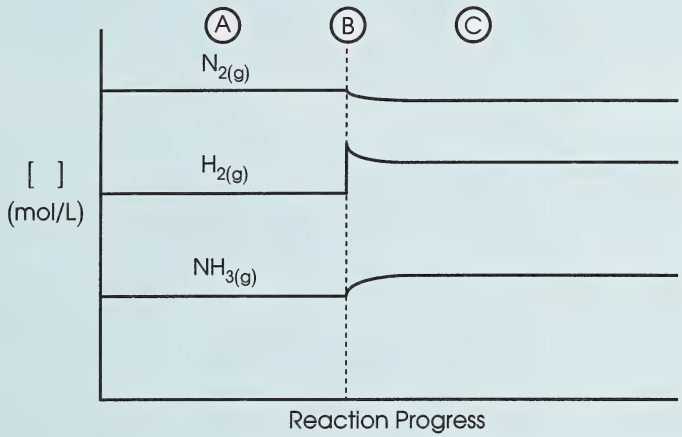
13. Choose the appropriate term from each set of brackets to fill in the blanks.

When extra reactants are added to the ammonia system at equilibrium, the system will relieve the stress by _____ (producing/decomposing) ammonia. When extra ammonia is added to the system, the stress is relieved by _____ (producing/decomposing) ammonia. The equilibrium shifts to _____ (use up/produce) the added substance.

Check your answers by turning to the Appendix, Section 1: Activity 3.

Concentration Change Effects on the Nitrogen-Hydrogen-Ammonia System

$$N_{2(g)} + 3\,H_{2(g)} \rightleftarrows 2\,NH_{3(g)} + 92\ kJ$$

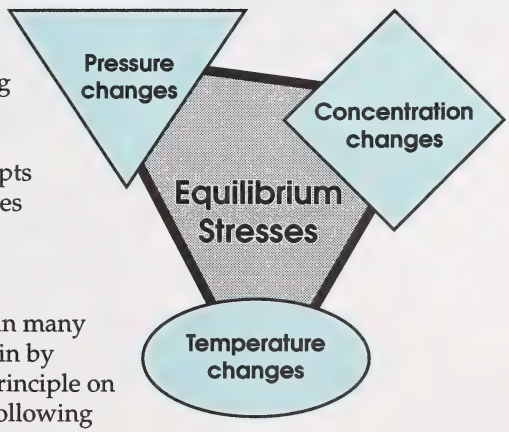


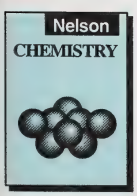
Ⓐ The system is at equilibrium.

Ⓑ An amount of $H_{2(g)}$ is added to the system. The concentration of $H_{2(g)}$ immediately increases.

Ⓒ The system tries to use up the extra $H_{2(g)}$ by shifting to produce more product. The $[N_{2(g)}]$ decreases, the $[H_{2(g)}]$ decreases, and the $[NH_{3(g)}]$ increases, proportional to their mole ratios. The system then comes to a new equilibrium.

You have seen from the preceding video and the examples given how chemical systems react to imposed stresses by shifting their equilibria to relieve the stress, as described in Le Châtelier's principle. The following pages will expand on these concepts to aid in your understanding of the principles involved.

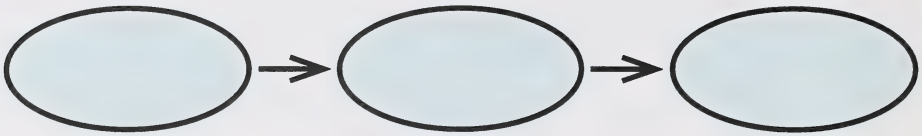




Your textbook also gives a description of Le Châtelier's principle and its importance in many industrial and technological processes. Begin by reading the introduction to Le Châtelier's principle on page 435 of your textbook and answer the following two questions.

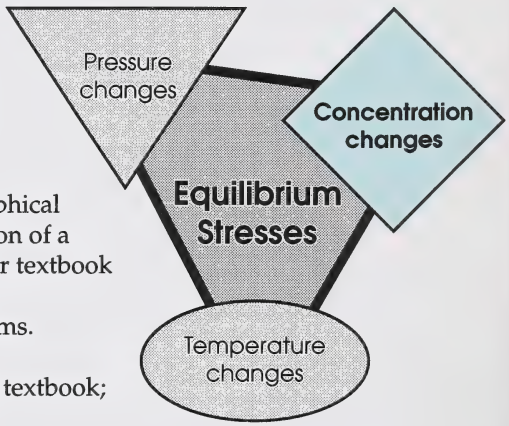
14. Why is controlling equilibrium states important technologically?

15. Complete the following organizer that summarizes the three stages evident when applying Le Châtelier's principle.

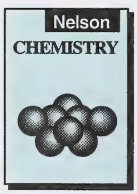


Check your answers by turning to the Appendix, Section 1: Activity 3.

Le Châtelier's principle is important because it allows you to predict how a chemical system will respond to applied stresses, such as changes in concentration, temperature, and pressure.



Earlier in this activity you were given a graphical example of how a change in the concentration of a reactant affects an equilibrium system. Your textbook also gives several examples of the effects of concentration changes on equilibrium systems.



Read the top two-thirds of page 436 in your textbook; then answer the next questions.

16. Complete the following table:

Stress	Direction Equilibrium Shifts
Add reactants	
Add products	

17. a. In one of the examples given on page 436 of your textbook, nitric acid is produced according to the following equation.

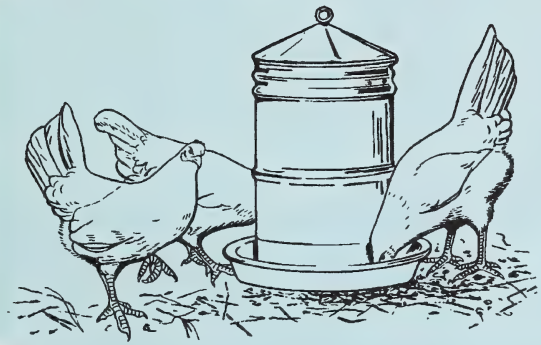
$$3\,NO_{2(g)} + H_2O_{(l)} \rightleftarrows 2\,HNO_{3(aq)} + NO_{(g)}$$

Besides the removal of nitrogen monoxide gas, what other concentration changes will cause a shift in the equilibrium towards the right?

b. Draw an equilibrium graph to show the effects on the equilibrium of increasing the $\left[NO_{2(g)}\right]$. Refer to the previous examples as a guide.

Check your answers by turning to the Appendix, Section 1: Activity 3.

DID YOU KNOW?



Chicken farmers may give their chickens carbonated water in hot temperatures to replenish the carbon dioxide content in the blood; the CO_2 is lost by panting to keep cool. The farmers are making use of Le Châtelier's principle because the presence of CO_2 forces production of the carbonate ion which in turn forces precipitation of $CaCO_3$, the main ingredient in egg shells. Lack of CO_2 causes thin shells and proportionately more broken shells.

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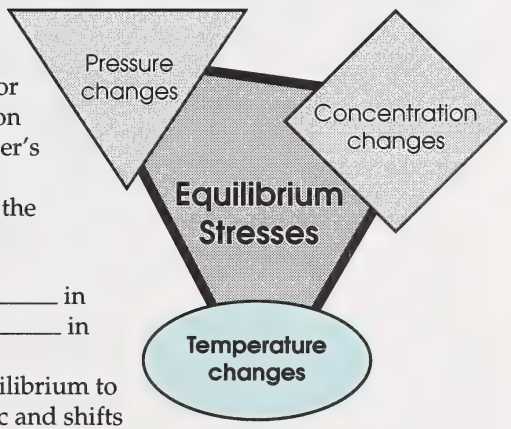
Energy can be treated as a reactant in endothermic reactions and as a product in exothermic reactions. As such, an increase or decrease in temperature will act as a stress on an equilibrium. Read the section Le Châtelier's Principle and Temperature Changes on pages 436-437 in your textbook and answer the following questions.

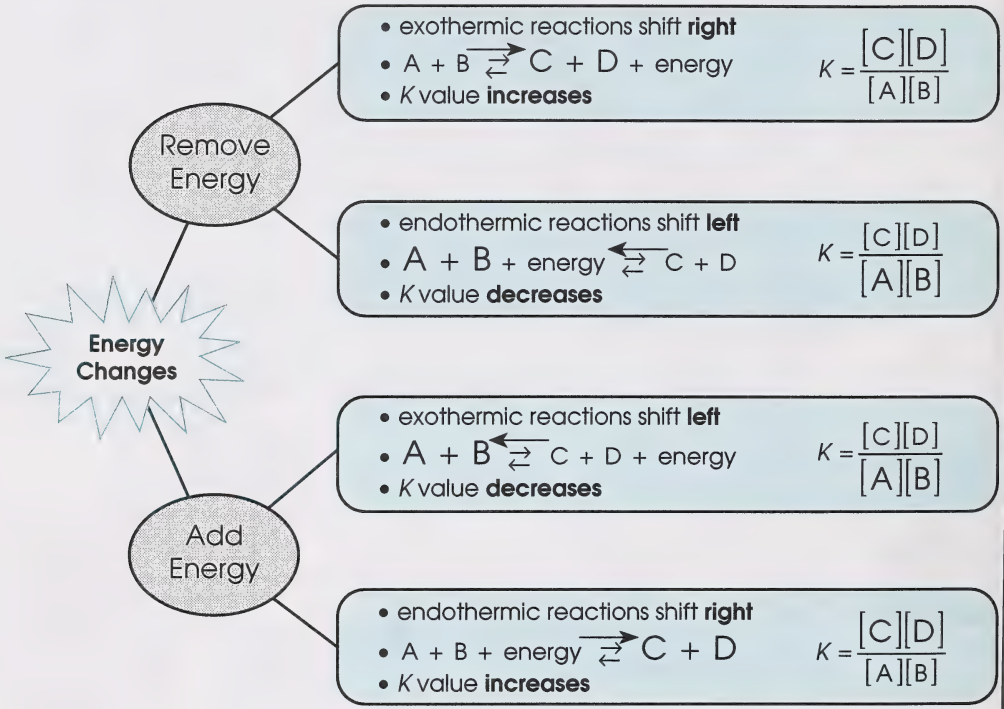


18. The addition of heat favours the _____ in an endothermic reaction and the _____ in an exothermic reaction. Lowering the temperature of a reaction shifts an equilibrium to _____ if the reaction is exothermic and shifts an equilibrium to reactants in an _____ process.

19. The following is a hypothetical reaction at equilibrium: $A + B \rightleftarrows C + D$

When heat is added to this system, the equilibrium shifts to the right. Is this an endothermic or exothermic reaction? Explain.

As predicted by Le Châtelier, the reaction always shifts to relieve the applied stress – in this case addition or removal of energy. However, since no other products or reactants have been added, the actual equilibrium constant (K) value will be different at new temperatures. *Very large changes in concentration and temperature changes are the only changes that affect the size of K.* An increase in temperature favours the direction that uses up energy; for an endothermic reaction the result is more products, less reactants, and hence a larger K value. The opposite is true for exothermic reactions. A decrease in temperature favours the direction that produces energy; for an exothermic reaction the result is more products and less reactants and hence a larger K value. Once again the opposite is true for endothermic reactions.

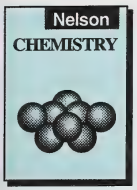

Energy Changes
Remove Energy
Add Energy
• exothermic reactions shift **right**
• A + B ⇄ C + D + energy
• *K* value **increases**
$K = \frac{[C][D]}{[A][B]}$
• endothermic reactions shift **left**
• A + B + energy ⇄ C + D
• *K* value **decreases**
$K = \frac{[C][D]}{[A][B]}$
• exothermic reactions shift **left**
• A + B ⇄ C + D + energy
• *K* value **decreases**
$K = \frac{[C][D]}{[A][B]}$
• endothermic reactions shift **right**
• A + B + energy ⇄ C + D
• *K* value **increases**
$K = \frac{[C][D]}{[A][B]}$

20. Consider the following system at equilibrium:

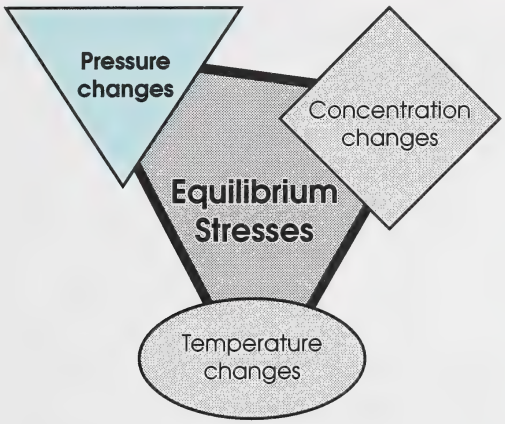
$$NaClO_{(aq)} + 2\,HCl_{(aq)} \rightleftarrows NaCl_{(aq)} + H_2O_{(l)} + Cl_{2(g)}$$

When this system is cooled, gas bubbles are produced. Is the system endothermic or exothermic? Explain.

Check your answers by turning to the Appendix, Section 1: Activity 3.



In an equilibrium system involving gases, the total pressure in the system can be changed by increasing or decreasing the volume of the system. To further clarify how this works, read Other Variables Affecting Chemical Systems on pages 437-438 in your textbook and answer the following questions.



21. List the terms necessary to fill in the blanks of the following paragraph.

Pressure and volume are inversely proportional; therefore, to increase the pressure on a system you ____a.____ the volume of the container and to decrease pressure you ____b.____ the volume. By Le Châtelier's principle, if the volume of a system is decreased, the equilibrium will shift to ____c.____ pressure by ____d.____ the total number of gas particles. If the volume is increased, the equilibrium will shift to ____e.____ pressure by ____f.____ the total number of gas particles. This shift is only productive if the total number of gas molecules on the reactant side in the equation is ____g.____ than the total number of gas molecules on the product side. An increase in pressure shifts the equilibrium to the side with the ____h.____ number of gas molecules; a decrease shifts the equilibrium to the side with the ____i.____ number of gas molecules.

22. How does a change in pressure affect a system with the same number of molecules on both sides of the equation?

23. a. How does the addition of a catalyst affect a system already at equilibrium?

b. How does the addition of a catalyst affect a system not yet at equilibrium?

Check your answers by turning to the Appendix, Section 1: Activity 3.

24. You have studied how changes in concentration, temperature, and pressure affect a system at equilibrium. Complete the following table to confirm your understanding of Le Châtelier's principle. In each case, describe which direction, if any, the equilibrium shifts and why.

Applied Stress	Response of System
Raise the temperature of an exothermic equilibrium system.	
Compress a gaseous system that has more reactant gases than product gases.	
Add a catalyst to a system at equilibrium.	
Add more of a reactant.	
Lower the temperature of an endothermic equilibrium system.	
Remove products from an equilibrium system.	
Add more solute to a saturated solution.	
Increase the volume of a gaseous equilibrium system.	
Add a catalyst to a system not yet at equilibrium.	

Check your answers by turning to the Appendix, Section 1: Activity 3.

Now that you are familiar with how Le Châtelier's principle is used to predict the way a system at equilibrium will change when conditions are changed, use what you have learned to do the following investigation.

Investigation: Equilibrium Shifts

Science Skills

- ☑ A. Initiating
- ☑ B. Collecting
- ☑ C. Organizing
- ☑ D. Analysing
- ☐ E. Synthesizing
- ☑ F. Evaluating

Carefully follow the directions for this investigation. Pay careful attention to the required components, safety aspects, and applied science skills.

The purpose of this investigation is to test Le Châtelier's principle by studying the following two chemical equilibrium systems:

① $Ca(OH)_{2(aq)} + CO_{2(g)} \rightleftarrows CaCO_{3(s)} + H_2O_{(l)}$

② $CO_{2(g)} + H_2O_{(l)} \rightleftarrows H^+_{(aq)} + HCO_3^-{}_{(aq)}$

The first equilibrium system is monitored by the presence of the precipitate that forms. The second equilibrium is commonly found in the human body and in soft drinks. Shifts in this equilibrium are detected by a diagnostic test involving the acid-base indicator bromothymol blue (HBb). Changes in the hydrogen ion concentration of this system can be detected by the colour changes shown by the indicator. HBb will turn blue with a decrease in the $\left[H^{+}_{(aq)}\right]$ to basic conditions and yellow (yellow-green) with an increase in $\left[H^{+}_{(aq)}\right]$ to acidic conditions. You will see both of these equilibrium systems in the same solution, which should lead to some interesting observations.

Problem

How does a change in the concentration of carbon dioxide affect Equilibrium ① ? How does a change in pressure affect Equilibrium ② ?

25. a. Predict how an increase in concentration of $CO_{2(g)}$ will affect Equilibrium ①.

    b. Predict how a change in pressure will affect Equilibrium ②.

Guidelines

You will prepare your own saturated $Ca(OH)_{2(aq)}$ (limewater) which will be used for this investigation, as well as other investigations in this module and Module 6. Prepare the limewater as follows:

- Set up the filtration apparatus as you did when you performed Investigation 14.1 in Activity 1 of this section.
- Fill a small test tube with about 8 mL of water.
- Add about 1 mL (1/4 tsp.) of $Ca(OH)_{2(s)}$ to the water.
- Stopper and shake the test tube for about one minute.
- Let the $Ca(OH)_{2(aq)}$ solution stand about five minutes.
- Filter the saturated $Ca(OH)_{2(aq)}$ solution into a graduated cylinder or other container.
- Store the limewater in a sealed, labelled container for use in this lab and subsequent labs in this course.

Materials

- 3 mL cold, fresh, colourless, carbonated water (soda water)
- 2 mL limewater (saturated $Ca(OH)_{2(aq)}$)
- 3 drops bromothymol blue indicator
- eyedropper
- small syringe (5 mL-50 mL) with no needle
- solid rubber stopper with partial hole to seal syringe tip
- 13 mm × 100 mm test tube
- small beaker or container
- lab apron
- safety glasses

Procedure

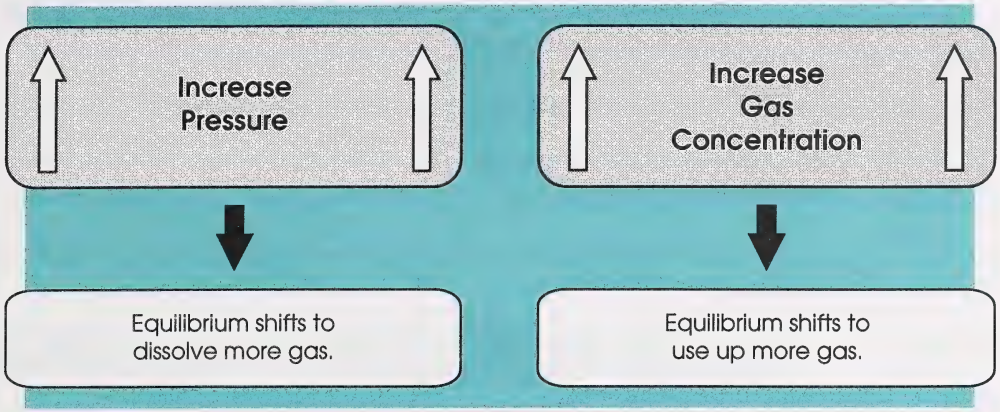
- Pour about 2 mL of limewater into a test tube.
- Add 2-3 drops of bromothymol blue indicator to the limewater and swirl to mix.
- With an eyedropper, add the soda water drop by drop to the limewater mixture. Observe any changes that occur.
- Stop adding soda water when the colour of the mixture has just turned yellow (yellow-green).
- Pour the limewater/soda water mixture into a small beaker or container.
- Draw about 2 mL of the limewater/soda water mixture into a syringe.
- Seal the tip of the syringe with a rubber stopper.
- Draw the syringe plunger up (back) in a series of about 2 mL-increments. Stop after each, swirl the solution, and observe any changes.
- Reverse the process with the syringe plunger once the top of the syringe has been reached.
- Make careful observations once the plunger is back to its original position.

26. Follow the procedure provided and record your evidence.

27. Did your results confirm Le Châtelier's principle and your predictions? Explain.

28. Evaluate the investigation. What variables were kept constant when the limewater was added? What variables were kept constant when the pressure was changed?

Check your answers by turning to the Appendix, Section 1: Activity 3.


Increase Pressure
Increase Gas Concentration
Equilibrium shifts to dissolve more gas.
Equilibrium shifts to use up more gas.

In the preceding investigation, you studied two different equilibrium systems – in one you saw how a change in concentration affects the equilibrium position; in the other you saw how a change in pressure affects the equilibrium position. The following investigation tests an equilibrium system for the effects of both temperature and concentration changes.

Investigation 14.3: Testing Le Châtelier's Principle

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Carefully follow the directions given for Investigation 14.3 on page 440 of your textbook. Pay special attention to the required components, safety aspects, and applied science skills.

Science Skills

- ☑ A. Initiating
- ☑ B. Collecting
- ☑ C. Organizing
- ☑ D. Analysing
- ☐ E. Synthesizing
- ☐ F. Evaluating

Guidelines

- This investigation is done in an ethanol solution so that the concentration of water can be altered. (The concentration of water is constant in an aqueous solution.)
- If you are preparing your own cobalt(II) chloride alcohol solution, dissolve 2 g of cobalt(II) chloride hexahydrate $\left(CoCl_2 \bullet 6H_2O_{(s)}\right)$ in 100 mL of ethyl alcohol (ethanol). (Ethanol can be obtained from a drugstore or pharmacy as a 93% v/v rubbing alcohol compound.) Add *just* enough water to the blue solution to make it an intermediate purple (blue-pink) colour.

Caution

- Use extreme caution with the alcohol. It is extremely flammable. Keep it away from all heat sources.
- Common table salt can be used as the $NaCl_{(s)}$ required.
- You can use 0.10 mol/L $AgNO_{3(aq)}$ instead of solid $AgNO_3$; add the solution a few drops at a time.

29. Predict how each of the proposed changes in the experimental design should affect the cobalt equilibrium system.

30. Write a brief description in point form of your procedure.

31. Prepare an appropriate data table, perform the procedure, and record your evidence.

32. Compare your evidence with the predictions you made by completing a table similar to the one shown.

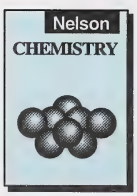
Applied Stress	Predicted Change	Observed Change
addition of heat		
reduction of heat		
addition of water		
addition of $NaCl_{(s)}$		
addition of $AgNO_{3(aq)}$		

33. Did Le Châtelier's principle hold true? Explain.

34. How does the addition of silver nitrate affect the equilibrium?

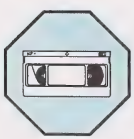
35. Evaluate the procedure followed. Are you confident of your results?

Check your answers by turning to the Appendix, Section 1: Activity 3.



36. Apply what you have learned about Le Châtelier's principle by doing textbook questions 12-16 from pages 440-441 in *Nelson Chemistry*.

Check your answers by turning to the Appendix, Section 1: Activity 3.



As mentioned earlier in the activity, Le Châtelier's principle is very important to chemical industries because it helps in optimizing conditions to make an industrial process economically viable. To study one such chemical process, the Haber process, read page 443 in your textbook and watch the program The Haber Process, from the video series *Chemical Equilibrium*. You may wish to preview the following questions before watching the video.

37. What international event prompted Haber to develop a process for synthesizing ammonia?

38. Using Le Châtelier's principle, what three stresses can be applied to the process to favour the production of ammonia?

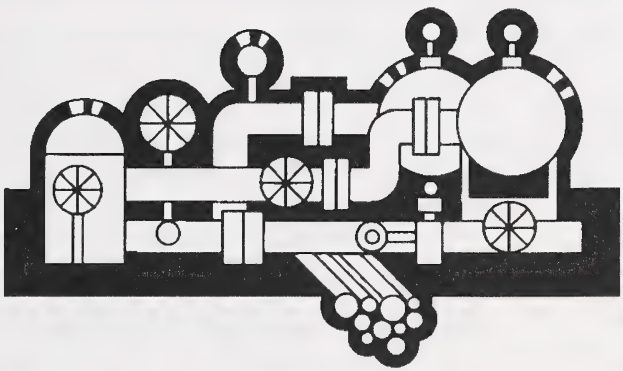
39. a. Why was a compromise needed when considering the effect of temperature?

    b. What was used to overcome the temperature problems?

40. Haber was recipient of the Nobel Prize in chemistry and members of the international community protested. Why was awarding Haber the prize contested?

Check your answers by turning to the Appendix, Section 1: Activity 3.

In this activity, you have learned to apply Le Châtelier's principle to predict how a system at equilibrium is affected when stressed and how to adjust chemical processes to make them easier to use. Using Le Châtelier's principle to help streamline industrial processes is an important economic application of chemistry.



Follow-up Activities

If you had difficulties understanding the concepts in the activities, it is recommended that you do the Extra Help. If you have a clear understanding of the concepts, it is recommended that you do the Enrichment.

Extra Help

Throughout this section, you have discovered that a system at equilibrium has constant observable properties and can be described qualitatively or measured quantitatively (by calculating percent reaction or the equilibrium constant). You have also seen that the equilibrium can be manipulated by changing the concentrations of both products and reactants, or changing the temperature, or changing the pressure at which the reaction is performed. The following questions will help you review the concepts studied.

1. Match the following terms with the descriptors.

Terms	Descriptors
a. catalyst	i. completely consumed in a quantitative reaction
b. equilibrium state	ii. $K \approx 1$
c. exothermic	iii. a mathematical relationship true at equilibrium
d. quantitative reaction	iv. forward rate = reverse rate
e. empirical properties	v. $K \gg 1$ ($\gg$ means *much greater than*)
f. non-spontaneous	vi. has no effect on the equilibrium conditions
g. percent reaction	vii. a process that releases heat
h. $\overset{50\%}{\rightleftarrows}$	viii. $K \ll 1$ ($\ll$ means *much less than*)
i. equilibrium law	ix. remain unchanged at equilibrium
j. limiting reagent	x. $\frac{\text{equilibrium } [\text{product}]}{\text{predicted } [\text{product}]} \times 100$

2. You are given the following chemical system at equilibrium:

$$CO_{2(g)} + H_{2(g)} \rightleftarrows CO_{(g)} + H_2O_{(g)} \qquad \Delta H_r = +42 \text{ kJ}$$

Use a chart similar to the following to describe what will happen to the equilibrium and the concentration of carbon monoxide when the various stresses are applied.

Stress	Direction of Shift	Change in $[CO_{(g)}]$
Decrease the volume of the container.		
Add heat.		
Remove $H_2O_{(g)}$ using a desiccator.		
Add $CO_{2(g)}$.		
Put the reaction vessel in an ice bath.		
Add a catalyst.		

desiccator – something which dries (removes water from) substances

3. You are given the following reaction:

$$CO_{2(g)} + H_{2(g)} \rightleftarrows CO_{(g)} + H_2O_{(g)}$$

At 427°C, the equilibrium constant (K) for the reaction is 0.20.

a. Write the equilibrium law for this reaction.

b. Use the equilibrium law to complete the following table.

Reaction	$[CO_{2(g)}]$ (mol/L)	$[H_{2(g)}]$ (mol/L)	$[CO_{(g)}]$ (mol/L)	$[H_2O_{(g)}]$ (mol/L)
1	1.8	2.0		1.2
2	0.5		0.3	2.0
3	8.0	14	35	

4. For the reaction, $2\ HI_{(g)} \rightleftarrows H_{2(g)} + I_{2(g)}$, 4.0 mol of colourless $HI_{(g)}$ was introduced into a 1.0 L container. The container was heated and then held at a constant temperature. A blue-pink colour started to appear, it intensified, and then it stayed constant. At this point, an analytical chemist determined that there was still 1.2 mol of $HI_{(g)}$ remaining in the container.

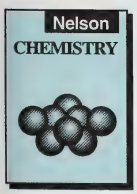
a. How do you know that equilibrium was reached?

b. Calculate the percent reaction.

c. Classify the reaction (favours products, favours reactants, or is quantitative).

d. Calculate the equilibrium constant (K) for the reaction.

Check your answers by turning to the Appendix, Section 1: Extra Help.



If you would like further practice in working with the concepts presented in this section, answer any of questions 1-7, 9, 10, 14-16, 18, and 19 on pages 456-457 of your textbook. The answers to these questions can be found on pages 518-519 of *Nelson Chemistry*.

Enrichment

Science Skills
- ☑ A. Initiating
- ☑ B. Collecting
- ☑ C. Organizing
- ☑ D. Analysing
- ☐ E. Synthesizing
- ☑ F. Evaluating

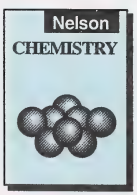
Do **one** of the following.

1. The extent to which an ionic compound is soluble in water can be measured by an equilibrium constant called the solubility product (K_{sp}). The K_{sp} of an ionic compound is calculated by multiplying the equilibrium concentrations of the ions raised to the power of their coefficients, which come from the dissociation equations for the ionic compounds.

> Example
>
> $$Ca(OH)_{2(s)} \rightleftarrows Ca^{2+}_{(aq)} + 2\ OH^-_{(aq)}$$
>
> $$K_{sp} = \left[Ca^{2+}_{(aq)}\right]\left[OH^-_{(aq)}\right]^2$$
>
> **Note:** The concentration of $Ca(OH)_{2(s)}$ is constant so a $Ca(OH)_{2(s)}$ term is not included in the equation for K_{sp}.

Design and carry out an investigation to determine the K_{sp} of $Ca(OH)_{2(s)}$. (**Hint:** From your knowledge gained in Chemistry 20, how can you find the concentration of a base such as $Ca(OH)_{2(s)}$?) Assume that the indicator endpoint is at pH = 7.

2. Visit or phone a local chemical plant (gas, ammonia, pharmaceutical, distillery) and ask what conditions are necessary to make the process that occurs at the location work to obtain maximum amounts of product. Most industrial chemical processes optimize conditions by applying some form of Le Châtelier's principle. What balance is achieved between economics and production?



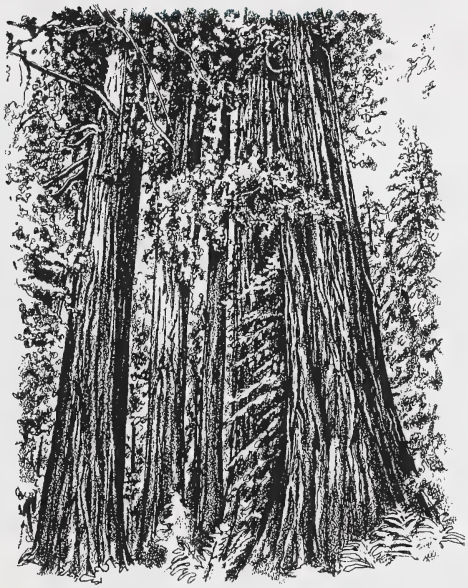
3. Do problems 35 and 36 in the section Extensions on page 458 of your textbook.

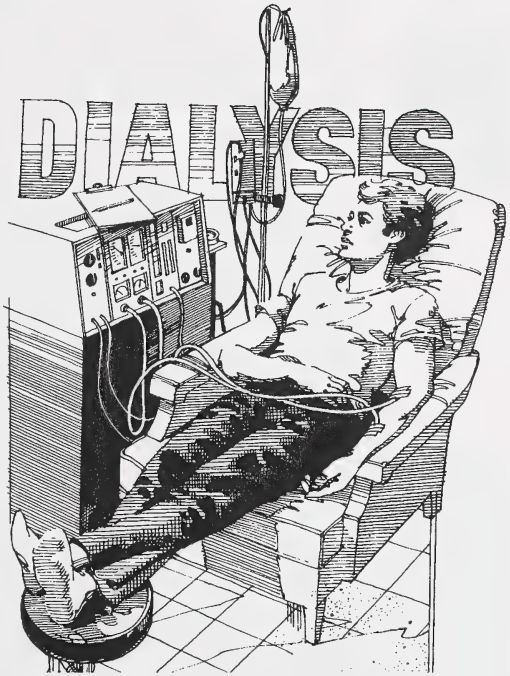
4. If you have access to supervised laboratory facilities, do Investigation 14.5: Studying a Chemical Equilibrium System on page 459 of your textbook. Include all parts of the investigation as indicated in your textbook. Carefully follow the directions given for the investigation. Pay special attention to the required components, safety aspects, and applied science skills.

Science Skills
- ☑ A. Initiating
- ☑ B. Collecting
- ☑ C. Organizing
- ☑ D. Analysing
- ☐ E. Synthesizing
- ☑ F. Evaluating

Check your answers by turning to the Appendix, Section 1: Enrichment.

Conclusion





In the preceding activities, you have discovered how to recognize, measure the extent of, and manipulate chemical equilibrium systems. These abilities will help you to understand such equilibria as those found in biological organisms. Biological systems have many interrelated equilibria and it is important to be able to understand how they can affect each other. Because water is such a common solvent, it is involved in many biological and other equilibrium systems. In the next section, you will examine the water equilibrium and how it affects solutions and other equilibria.

30

Section 1 Assignment: Systems in Balance

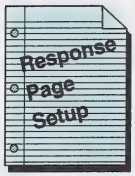
Review the Evaluation information found in the introductory pages of this module.

It is important to number and clearly identify each page with the following information at the top:

Chemistry 30 – Module 5 Section 1 Assignment Page # Name and ID#

Be sure to write legibly. Leave a wide left margin and number all of your pages.

Part A: Multiple Choice



Read each question carefully and decide which of the choices **best** answers the question or completes the statement. Place the number of each question in a vertical column and place your answer (A, B, C, or D) beside each number. **(Total = 5 marks)**

Use the following information to answer questions 1-3.

I. $SO_{2(g)} + Cl_{2(g)} \rightleftarrows SO_2Cl_{2(g)} + \text{energy}$

II. $CH_{4(g)} + H_2O_{(g)} + \text{energy} \rightleftarrows CO_{(g)} + 3\,H_{2(g)}$

III. $2\,SO_{2(g)} + O_{2(g)} \rightleftarrows 2\,SO_{3(g)} + \text{energy}$

IV. $3\,O_{2(g)} + \text{energy} \rightleftarrows 2\,O_{3(g)}$

V. $H_{2(g)} + I_{2(g)} + \text{energy} \rightleftarrows 2\,HI_{(g)}$

1. Which of the system(s) would shift towards the products with the addition of heat?

   A. I, III
   B. II
   C. II, IV, V
   D. V

2. The system(s) which would shift towards reactants with an increase in pressure by compressing the system into a smaller vessel would be

   A. I, III, IV
   B. II
   C. IV
   D. V

3. The system whose equilibrium would not shift because of any change in pressure is

   A. I
   B. III
   C. IV
   D. V

4. Which of the following statements is true about a chemical system at equilibrium?

   A. $K = 1$
   B. [products]/[reactants] = 1
   C. [reactants] = [products]
   D. forward reaction rate = reverse reaction rate

5. A particular chemical equilibrium has an equilibrium constant $K = 1250$. A true statement about this equilibrium would be

   A. The products are favoured.
   B. The reactants are favoured.
   C. The reaction is exothermic.
   D. The reaction reaches equilibrium quickly.

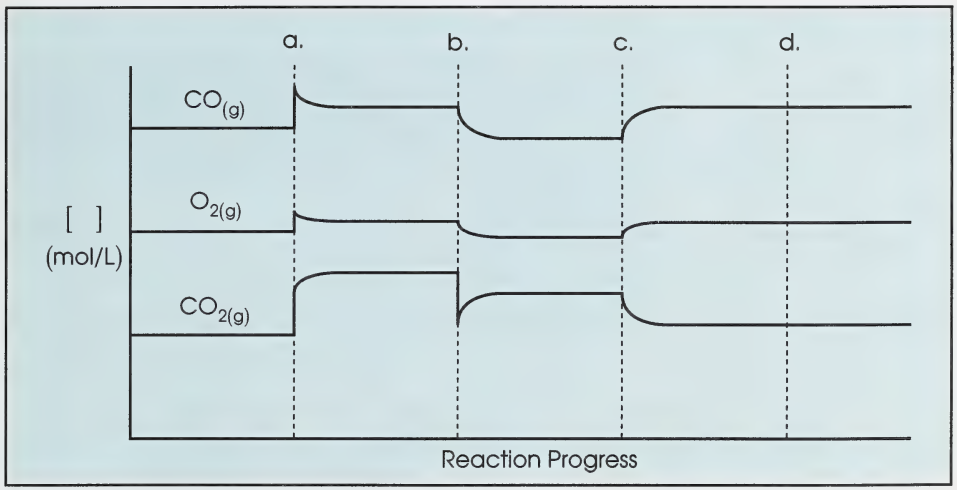
Part B: Written Response

Response Page Setup

Write your answers neatly and concisely. For full marks, your answers must show all pertinent explanations, calculations, and formulas.

6. The following is a graphical representation of this equilibrium:

$$2\,CO_{(g)} + O_{2(g)} \rightleftarrows 2\,CO_{2(g)} + \text{energy}$$



Describe what stress must have occurred to create each situation shown on the graph. The stresses are marked by a., b., c., and d. **(4 marks)**

7. In a closed environment, phosphorus pentachloride will decompose into, and form an equilibrium system with, phosphorus trichloride and chlorine gas.

$$PCl_{5(g)} \rightleftarrows PCl_{3(g)} + Cl_{2(g)}$$

Calculate the equilibrium constant for this reaction at 25°C if a set of equilibrium concentrations were found to be $\left[PCl_{3(g)}\right] = \left[Cl_{2(g)}\right] = 0.850$ mol/L and $\left[PCl_{5(g)}\right] = 0.406$ mol/L. **(2 marks)**

8. In a flask initially containing colourless molecules of $N_2O_{4(g)}$, the brown colour characteristic of $NO_{2(g)}$ appears and the colour steadily intensifies. Is this system in equilibrium? Explain. **(2 marks)**

9. The following reaction was performed at 1000°C.

$$N_{2(g)} + 3\,H_{2(g)} \rightleftarrows 2\,NH_{3(g)} \qquad \Delta H = -92\ kJ$$

The equilibrium constant was found to be 0.0232 (Reaction 1). When this same reaction was performed at a different temperature, the equilibrium constant was calculated as 0.0563 (Reaction 2).

Was the temperature in Reaction 2 higher or lower than 1000°C? Explain your reasoning. **(3 marks)**

10. **EQUILIBRIUM CONDITIONS FOR NO – O_2 – NO_2 SYSTEM AT 227°C**

System	$[NO_{(g)}]$ (mol/L)	$[O_{2(g)}]$ (mol/L)	$[NO_{2(g)}]$ (mol/L)
1	2.50×10^{-2}	5.58×10^{-8}	1.50×10^{-1}
2	1.50×10^{-2}	i	2.00×10^{-1}
3	ii	3.00×10^{-3}	7.50×10^{-2}
4	2.00×10^{-4}	1.00×10^{-5}	iii

a. Write the equilibrium law expression and use the preceding table to calculate the K value for the equilibrium system $2\,NO_{(g)} + O_{2(g)} \rightleftarrows 2\,NO_{2(g)}$ at 227°C. **(2 marks)**

b. Calculate the values necessary to complete the rest of the table (i, ii, and iii). Show all your work. **(3 marks)**

11. Dinitrogen tetraoxide exists in an equilibrium with nitrogen dioxide:

$$N_2O_{4(g)} \rightleftarrows 2\,NO_{2(g)}$$

An amount of 0.40 mol of $N_2O_{4(g)}$ was placed in a 1.0 L container which was heated to 100°C. When equilibrium was reached, the concentration of $N_2O_{4(g)}$ was found to be 0.16 mol/L.

a. How many moles of $N_2O_{4(g)}$ were used up in reaching equilibrium? **(1 mark)**

b. How many moles of $NO_{2(g)}$ were produced and what was the $\left[NO_{2(g)}\right]$ at equilibrium? **(2 marks)**

c. Write the equilibrium law expression for this system. **(1 mark)**

d. Calculate the equilibrium constant for this system at 100°C. **(1 mark)**

e. What macroscopic property could you observe in this system to establish that equilibrium was reached? **(1 mark)**

Science Skills

- ☐ A. Initiating
- ☐ B. Collecting
- ☐ C. Organizing
- ☑ D. Analysing
- ☐ E. Synthesizing
- ☐ F. Evaluating

12. As a project in Chemistry 30, Pat was assigned the task of proving that equilibrium is dynamic. The teacher provided Pat with a saturated solution of lead(II) sulphate in a test tube (some solid lead(II) sulphate remained on the bottom of the test tube), a small sample of solid lead(II) sulphate which contained radioactive ^{35}S, and a stopper that fit the test tube. Pat placed the radioactive sample in the test tube containing the saturated solution of lead(II) sulphate, stoppered the test tube, and agitated it vigorously for several minutes. Pat then let the mixture settle overnight, filtered it, and tested both the filtrate and the residue for radioactivity.

Evidence: Both the residue and the filtrate gave a positive test for radioactivity. Also, the mass of the residue remained constant.

Complete the analysis for this investigation. How does the investigation show that equilibrium is dynamic? **(3 marks)**

Response Page Setup

Reproduce the science skills assessment box after your response to this question. Remember to indicate your evaluation of your skill level for each identified science skill. Leave a space for the teacher assessment.

Self:	A.	☐	B.	☐	C.	☐	D.	☐	E.	☐	F.	☐
Teacher:	A.	☐	B.	☐	C.	☐	D.	☐	E.	☐	F.	☐

SECTION 2

Water's Balance of Power

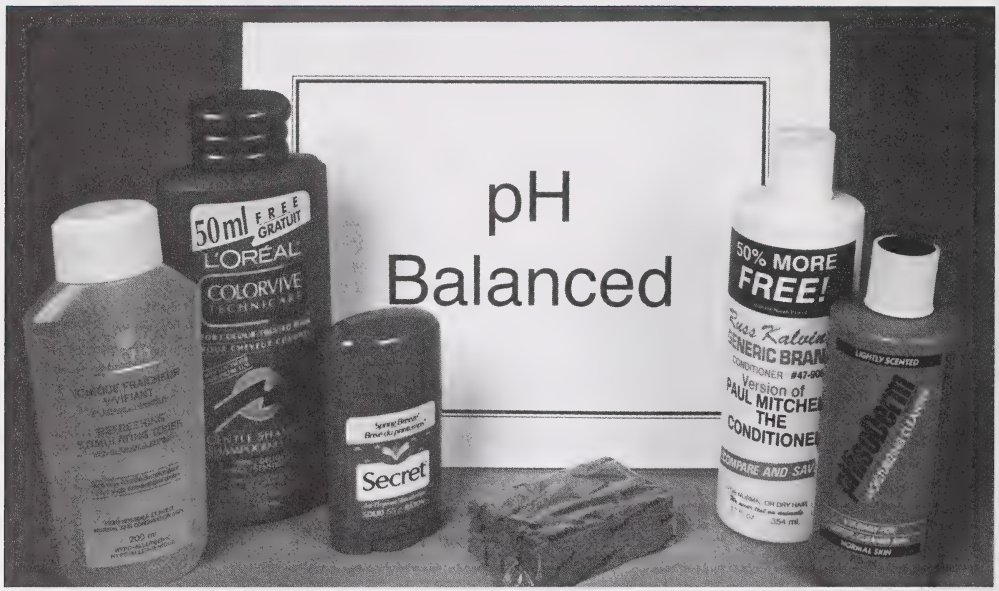


PHOTO SEARCH LTD.

You often see the advertising catchphrase *pH balanced* for a myriad of products you are being persuaded to buy. But what does the term *pH* mean and what is it balanced against? These are questions that you as a consumer should be asking yourself before purchasing any of the products.

In this section, you will be continuing your study of balances in systems, specifically in that very important system – water! You will examine the specific balance that exists in water and how pH is related to that balance. You will look at the concepts of acid-base strengths and weaknesses. Finally, you will explore a way to apply the equilibrium constant, *K*, to acids. As a consumer, this knowledge may help you become more discerning when reading or viewing advertisements.

ACTIVITY 1

Equilibrium in Water

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Most people think of acids and bases as being dangerous chemicals that should be handled with care (this is true in many cases). However, many of the everyday substances that you use for cleaning or eat or drink are acidic or basic in nature. What makes things acidic or basic? Empirical properties can be used to tell you if something is an acid or base, but why is it an acid or base? That is another matter.

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In Chemistry 20, you studied the empirical properties of acidic, basic, and neutral solutions. Before examining the reasons why these solutions exhibit different properties it is necessary that you review what those empirical properties are. Review page 114 in your textbook and then answer the following questions.

1. Why is measuring the conductivity of a solution not a suitable diagnostic test for distinguishing between an acidic, basic, and neutral solution?

2. What is a suitable diagnostic test to distinguish between acidic, basic, and neutral solutions?

Check your answers by turning to the Appendix, Section 2: Activity 1.

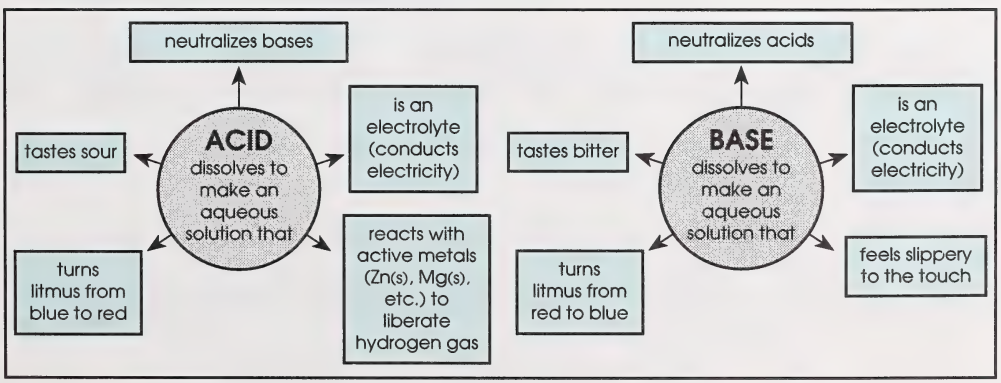
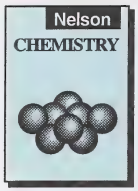

neutralizes bases
tastes sour
ACID
dissolves to make an aqueous solution that
is an electrolyte (conducts electricity)
turns litmus from blue to red
reacts with active metals (Zn(s), Mg(s), etc.) to liberate hydrogen gas
neutralizes acids
tastes bitter
BASE
dissolves to make an aqueous solution that
is an electrolyte (conducts electricity)
turns litmus from red to blue
feels slippery to the touch

FIGURE 5.1: Acids and Bases

The title of this activity is Equilibrium in Water. What does equilibrium in water have to do with acids and bases? How does knowing this relationship answer the earlier question that was asked, *why* is it an acid or a base?



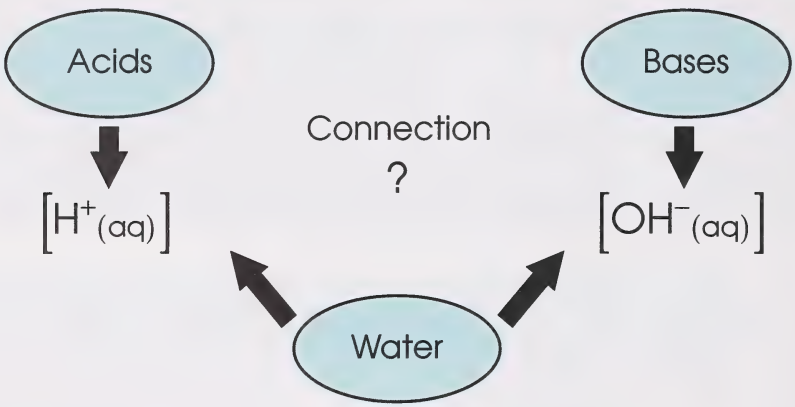
These are the types of questions that Svante Arrhenius was trying to answer in his studies in 1887. To rediscover (this material was covered in Chemistry 20) his theoretical ideas about solutions, acids, and bases, read pages 115-117 (omitting the biography on page 116) of your textbook.

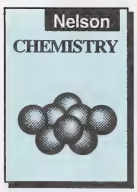
3. Compare the terms *dissociation* and *ionization* with respect to solutes in solution.

4. Write a theoretical (Arrhenius) definition and an empirical definition of an acid. Write a theoretical (Arrhenius) definition and an empirical definition of a base.

5. Notice that in Figure 5.1, taste is included as an empirical property. Is tasting a suitable method for classifying substances? Explain your answer. Why is it included?

Check your answers by turning to the Appendix, Section 2: Activity 1.





Do you begin to see a connection between an equilibrium involving water, and acids and bases? For further clues, read page 442 and the top of page 443 in *Nelson Chemistry*.

6. a. Using the Arrhenius theory, what does the presence of conductivity in pure water tell you?

   b. What does the relative degree of the conductivity tell you?

   c. Write the equation for the dissociation of water into ions, including percent reaction.

   d. How would you classify this water equilibrium?

K_w – equilibrium constant for the ionization of water

As with other equilibrium systems, an equilibrium constant can be found for the water system – the ionization constant for water, K_w.

7. a. Write the expression for the ionization constant of water.

   b. Why is $[H_2O_{(l)}]$ left out of this expression?

8. a. How does the size of K_w relate to the extent of the equilibrium?

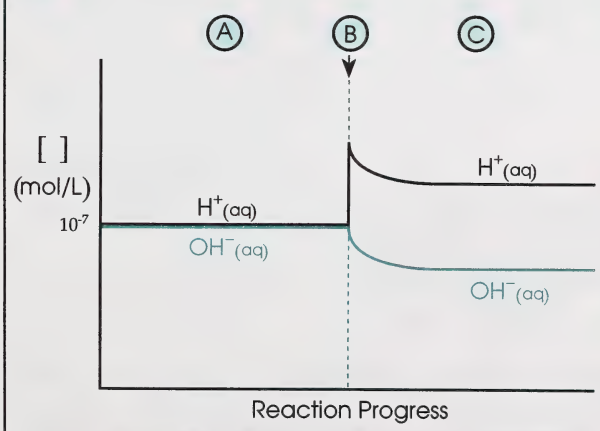
   b. Does this size of K_w agree with the conductivity evidence for pure water?

Check your answers by turning to the Appendix, Section 2: Activity 1.

auto-ionization – ionization without external influences

Water undergoes **auto-ionization** as illustrated by the equilibrium equation you just wrote. Like other equilibria, when acted on by stresses, it follows Le Châtelier's principle. Since all aqueous solutions use water as a solvent, the equilibrium between water molecules, hydrogen ions, and hydroxide ions is particularly important. When $H^+_{(aq)}$ ions are introduced into an aqueous system (as in the dissolving and ionizing of an acid in water), the equilibrium shifts to the left, using up $OH^-_{(aq)}$ ions. Likewise, when $OH^-_{(aq)}$ ions are introduced into an aqueous system (as in the dissolving and dissociating of a base in water) the equilibrium also shifts to the left, but uses up $H^+_{(aq)}$ ions.

1 **Graph for Water Equilibrium** $H_2O_{(l)} \rightleftarrows H^+_{(aq)} + OH^-_{(aq)}$

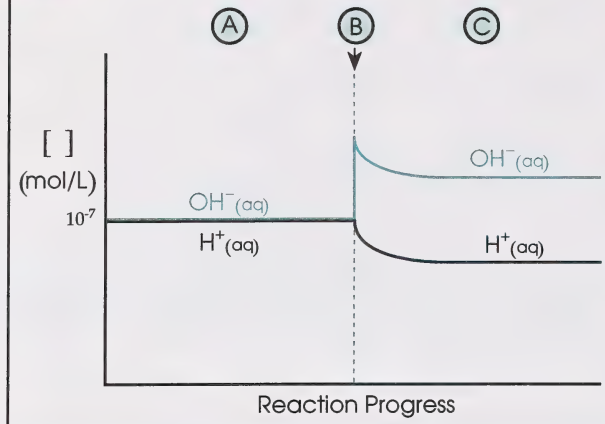


(A) Pure water at equilibrium: $[H^+_{(aq)}] = [OH^-_{(aq)}] = 10^{-7}$ mol/L

(B) Add an acid to the system – the $[H^+_{(aq)}]$ increases immediately.

(C) The equilibrium shifts left and comes to a new equilibrium where $[H^+_{(aq)}] > [OH^-_{(aq)}]$.

2 **Graph for Water Equilibrium $H_2O_{(l)} \rightleftarrows H^+_{(aq)} + OH^-_{(aq)}$**



(A) Pure water at equilibrium: $[H^+_{(aq)}] = [OH^-_{(aq)}] = 10^{-7}$ mol/L

(B) Add a base to the system – the $[OH^-_{(aq)}]$ increases immediately.

(C) The equilibrium shifts left and comes to a new equilibrium where $[OH^-_{(aq)}] > [H^+_{(aq)}]$.

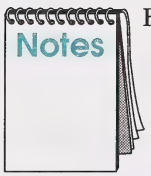
In Acid Solution

$K_w = 1.0 \times 10^{-14}\ (mol/L)^2 = [H^+(aq)][OH^-(aq)]$

In Base Solution

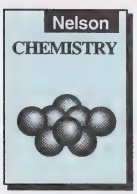
$K_w = 1.0 \times 10^{-14}\ (mol/L)^2 = [H^+(aq)][OH^-(aq)]$

at SATP



Here are some important points to remember:

- As long as there is water present (*any* aqueous solution), the water equilibrium, K_w, applies.
- No matter how strongly acidic or basic a solution is, there will *always* be some of each type of ion ($H^+_{(aq)}$ and $OH^-_{(aq)}$) present.
- The K_w value is constant at $1.0 \times 10^{-14}\ (mol/L)^2$ for any aqueous solution at SATP.
- The units for K_w are $(mol/L)^2$ and they are usually included in the equilibrium expression.



One application of the K_w equilibrium expression and value is to calculate the concentration of either the hydrogen ion or hydroxide ion if the other is known. Read through the calculation examples on page 444 in your textbook paying close attention to unit cancellation and significant digits.

Practise using the K_w expression to calculate concentrations by doing the following question.

9. Do textbook questions 17-22 on page 445 and textbook question 20 on page 457 of *Nelson Chemistry*.

Check your answers by turning to the Appendix, Section 2: Activity 1.

You will often find that there is more than one equilibrium present in a solution. A change in one of these equilibria may cause a change in another equilibria.

Many aqueous equilibria are affected by the addition of acids and bases. These equilibria are acting together with the water equilibrium. One such equilibrium is the chromate-dichromate equilibrium.

Lab Exercise 14C: The Chromate-Dichromate Equilibrium

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10. Complete Lab Exercise 14C on page 445 of your textbook. Pay special attention to the required science skill.

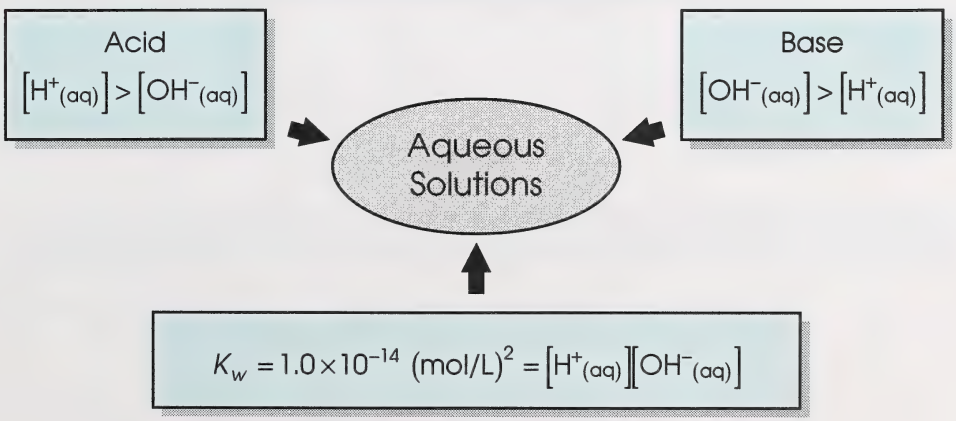
    a. Predict how changing the $[H^+_{(aq)}]$ will affect the chromate-dichromate equilibrium.

    b. Design an investigation to test your prediction.

Science Skills

- [x] A. Initiating
- [ ] B. Collecting
- [ ] C. Organizing
- [ ] D. Analysing
- [ ] E. Synthesizing
- [ ] F. Evaluating

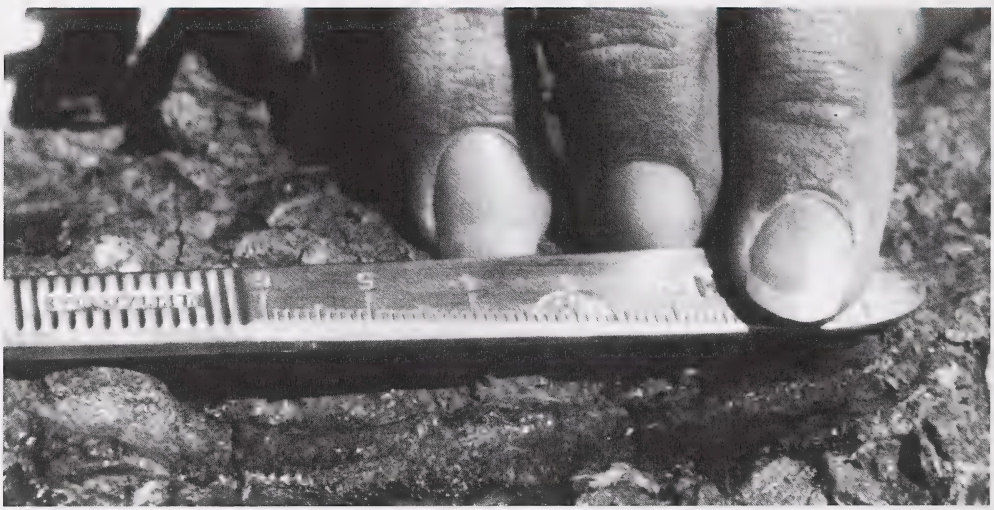
Check your answers by turning to the Appendix, Section 2: Activity 1.



In this activity you have discovered that substances that demonstrate the empirical properties of acids have a high $[H^+_{(aq)}]$, bases have a high $[OH^-_{(aq)}]$, and that these two ions are produced in the water equilibrium that exists in all aqueous solutions. However, the $[H^+_{(aq)}]$ and $[OH^-_{(aq)}]$ are often so small that expressing their concentrations can be inconvenient. In the next activity you will learn about a different scale used to communicate the acidity and basicity of a solution.

ACTIVITY 2

The Power of Hydrogen



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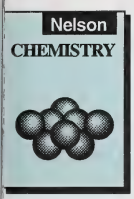
How do you measure things that are very small? very large? How do you measure quantities of length? of temperature? There are a variety of measurement scales for measuring various characteristics.

Some measurement scales have to be able to measure over a wide range. A scale used to measure loudness has to be able to measure both a quiet whisper and the roar of jet engines. The decibel scale measures all ranges of sound, and is based on powers of 10 (a logarithmic scale). The Richter scale measures all sizes of earthquakes and is also logarithmic. In this activity, you will learn to use a similar scale to measure acidity and basicity.

In Chemistry 20, you studied the pH scale and its relationship to acidity and basicity. A low pH indicated an acidic solution, a high pH indicated a basic solution, and a pH of 7 indicated a neutral solution. You dealt with whole number pH values and found that the pH corresponded to the negative of the power of 10 of the $[H^+_{(aq)}]$.

$$pH = -\log[H^+_{(aq)}]$$

In this course you will see that most pH values you will work with are not whole numbers. This necessitates a need for understanding how pH is calculated.

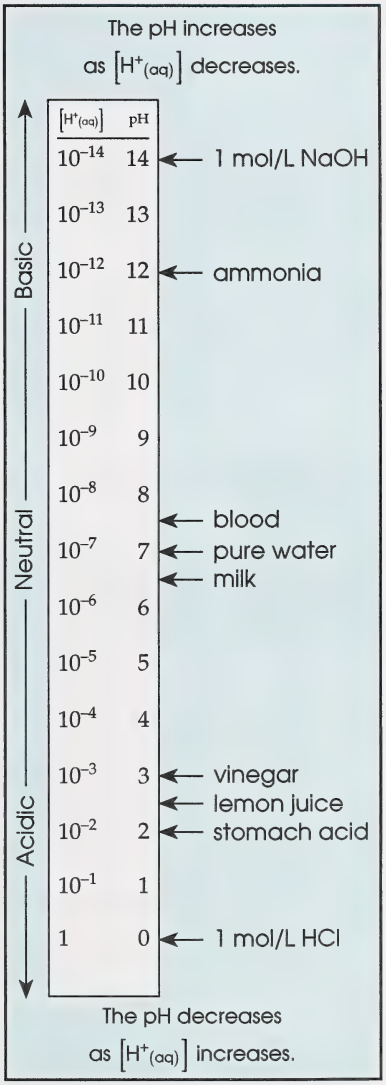


To learn how the pH scale is calculated, read pages 445-446 in your textbook, paying careful attention to the sample calculations and calculator instructions in the margin. Your calculator may not work in the manner noted; you may have to refer to the instruction manual for your calculator.

Note: If you become familiar with all applicable functions on your calculator now, it will save you time later when doing assignments and writing your final examination.

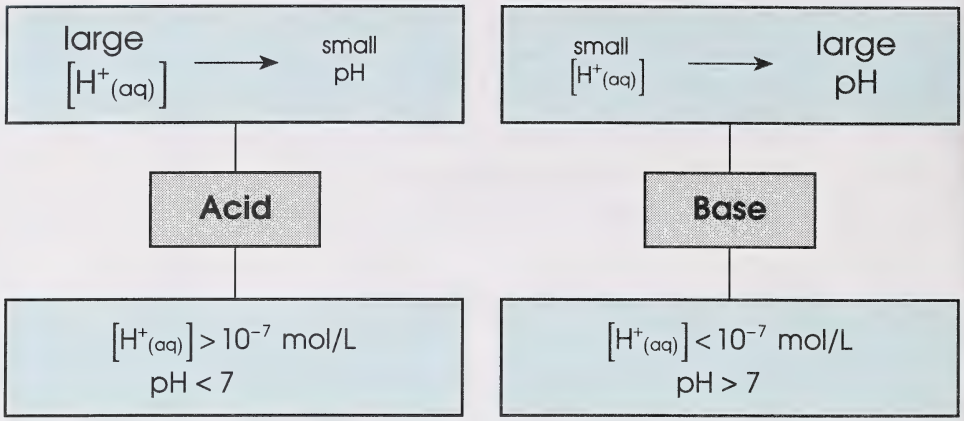
1. The number of significant digits communicated in the hydrogen ion concentration is the same as _____.

2. If $pH = -\log[H^+_{(aq)}]$, then $[H^+_{(aq)}]$ = _____.



3. Complete the following table of pH values, $[H^+_{(aq)}]$, and acidic-basic classification.

pH	$[H^+_{(aq)}]$ (mol/L)	Classification (acidic/basic/neutral)
3.67		
	2.45×10^{-5}	
10.1		
	9.345×10^{-12}	
5.63		



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From Activity 1, you know that $K_w = [H^+(aq)][OH^-(aq)] = 1.0 \times 10^{-14}$ $(mol/L)^2$. Just as pH can be calculated from $[H^+(aq)]$, pOH can be calculated from $[OH^-(aq)]$. Read pages 446-447 of your textbook to learn about pOH and how it relates to pH; then, answer the following questions.

4. Two very useful relationships to use when converting between pH, pOH, $[H^+(aq)]$, and $[OH^-(aq)]$ are

   - pH + pOH = _____ at SATP
   - _____ × _____ = 1.0×10^{-14} $(mol/L)^2$ at SATP

5. Do textbook question 23 on page 447 of *Nelson Chemistry*.

6. You are given the following question:

   > Calculate the pH of a solution made by dissolving 2.4 g of $NaOH_{(s)}$ in enough water to make 250 mL of solution.

   a. Indicate, in words, the steps you would perform to answer the question.

   b. Find the answer to the question.

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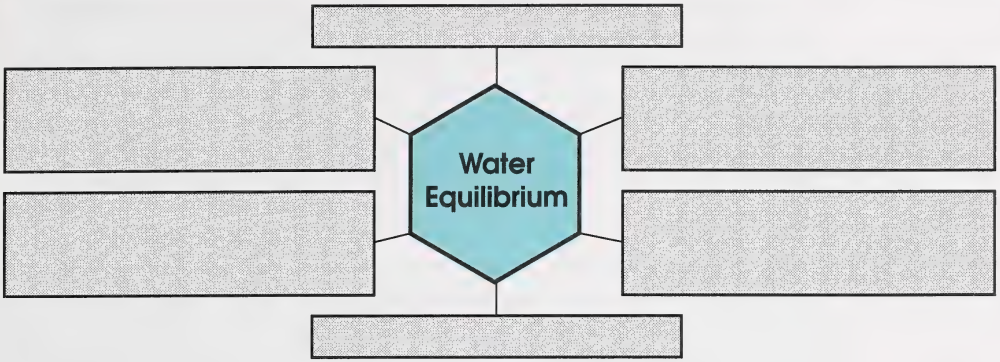
7. Do textbook questions 24 and 25 on page 447 and textbook question 29 on page 458 of *Nelson Chemistry*.

Nelson CHEMISTRY

8. Summarize the various water equilibrium calculations you can perform (as shown on page 447 of your textbook) in a chart similar to the following:



9. Complete the following table comparing pH, $[H^+_{(aq)}]$, volume, and mass of solute for $HCl_{(aq)}$ solutions.

pH	$[H^+_{(aq)}]$ (mol/L)	Volume of Solution	Mass of HCl Needed
3.50		250 mL	
	2.45×10^{-3}	2.00 L	
		400 mL	3.65 g
2.785		3.50 L	
3.00			7.3 mg

10. Complete the following table comparing pOH, pH, $[OH^-_{(aq)}]$, volume, and mass of solute for $NaOH_{(aq)}$ solutions. Note that when a solution is basic, it is imperative to use $[OH^-_{(aq)}]$ in the calculation, not $[H^+_{(aq)}]$.

pH	pOH	$[OH^-_{(aq)}]$ (mol/L)	Volume	Mass of NaOH
10.35			250 mL	
	2.40			24 g
		3.5×10^{-5}		2.00 g
		6×10^{-1}	3 L	
	1.455			36.00 mg

Check your answers by turning to the Appendix, Section 2: Activity 2.

How do changes in pH relate to changes in $[H^+_{(aq)}]$? If you double the pH of a solution does the $[H^+_{(aq)}]$ also double? Because pH and pOH are in a logarithmic scale they respond differently from other mathematical relationships.

$$pH - 1 = 10 \times [H^+_{(aq)}] \quad (10^1 \times [H^+_{(aq)}])$$
$$pH - 2 = 10 \times 10 \times [H^+_{(aq)}] \quad (10^2 \times [H^+_{(aq)}])$$
$$pH + 1 = \frac{1}{10} \times [H^+_{(aq)}] \quad (10^{-1} \times [H^+_{(aq)}])$$

Example

If the pH of a lake changes one pH unit from 5.0 to 4.0, what is the corresponding change in $[H^+_{(aq)}]$?

Answer:

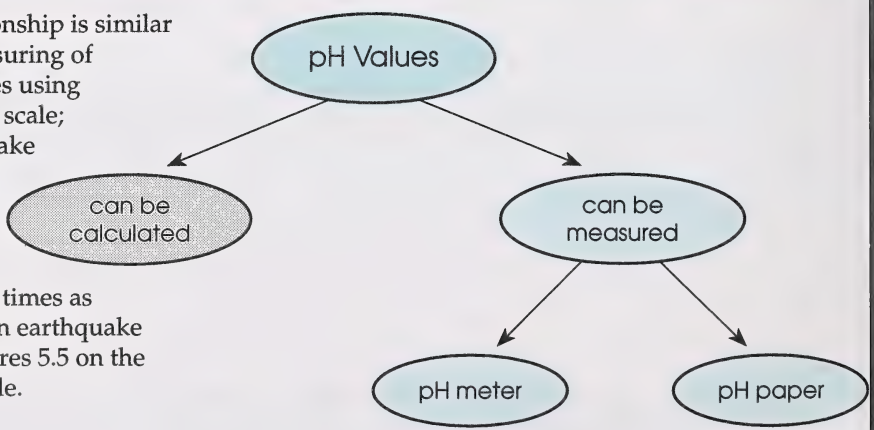
$pH = 5.0 \rightarrow \quad [H^+_{(aq)}] = 10^{-5.0} = 1 \times 10^{-5}$ mol/L

$pH = 4.0 \rightarrow \quad [H^+_{(aq)}] = 10^{-4.0} = 1 \times 10^{-4}$ mol/L

$5.0 - 4.0 = 1.0$ pH unit $\rightarrow \quad \frac{1 \times 10^{-4} \text{ mol/L}}{1 \times 10^{-5} \text{ mol/L}} = 10^1$ or tenfold increase in $[H^+_{(aq)}]$

This means that if the pH of a lake decreases by 1 pH unit (possibly due to acid rain), the $[H^+_{(aq)}]$ actually increases 10 times!

This relationship is similar to the measuring of earthquakes using the Richter scale; an earthquake measuring 6.5 on the Richter scale is actually 10 times as strong as an earthquake that measures 5.5 on the Richter scale.



You have seen how pH values can be calculated. The pH values can also be measured and then used to compare acidity or basicity of solutions such as lake water. There are two major ways to measure pH – using a pH meter or using pH paper.

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A pH meter makes use of electrode potentials like the ones you learned about in Module 4. The pH meter measures the change in potential difference caused by the solution between its reference electrode and its measuring electrode, and converts the change into a pH value. This value is usually displayed digitally or on a meter (analog) readout. A photograph of a pH meter is displayed in Figure 14.13 on page 447 of your textbook.

Read the section entitled The pH Meter at the bottom of page 447 and the top of page 448 of your textbook to prepare you for the investigation that follows.

Using pH paper is not as accurate as a pH meter, but it is sometimes easier to use. The pH paper changes colour depending on the pH of the solution it is in contact with, and you determine the pH by comparing the colour of the paper to a standard colour chart. In the investigation that follows, you will use either a pH meter or pH paper for determining the pH of common substances.

Investigation 14.4: pH of Common Substances

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Carefully follow the directions given for Investigation 14.4 on page 448 of your textbook. Pay special attention to the required components, safety aspects, and applied science skills.

Science Skills

- ☐ A. Initiating
- ☑ B. Collecting
- ☑ C. Organizing
- ☑ D. Analysing
- ☐ E. Synthesizing
- ☐ F. Evaluating

PATHWAYS

If you have access to supervised laboratory facilities, do Part A.
If you do not have access to laboratory facilities, do Part B.

Part A

Guidelines

buffer – a solution that maintains a near-constant, specified pH

- Check that your pH meter is working properly. Check with your teacher or learning facilitator if the meter does not read close to the pH of the buffer when tested.
- Follow the rinsing and calibrating instructions carefully.
- Dissolve each substance in water before testing with the pH meter; the electrode tip must be completely immersed in the solution.
- You should test at least seven different substances.

11. Follow the procedure outlined in the textbook and record your data in an appropriate evidence table.

Check your answers by turning to the Appendix, Section 2: Activity 2.

You will complete your analysis by doing the questions that follow Part B.

End of Part A

Part B

Guidelines

- You will use pH paper in place of a pH meter.
- Have a pH colour chart handy for comparison. If you are colour-blind, you may require someone else to compare the colours for you.
- Dissolve the substances in water before testing with the pH paper.
- You should test at least seven substances.

12. Perform your procedure and record your evidence in an appropriate evidence table.

Check your answers by turning to the Appendix, Section 2: Activity 2.

End of Part B

Complete your analysis by answering the question that follows.

13. Analyse the evidence you collected. Try and make a generalization about the types of substances you tested and their acidity or basicity.

Check your answers by turning to the Appendix, Section 2: Activity 2.

In this activity, you have discovered how to calculate the pH and pOH of solutions and to relate them to the acidity or basicity of different substances. How do these values relate to the strengths of acids or bases? How do they relate to the equilibrium position of water-based systems? These relationships will be explored in the final two activities of this module.

ACTIVITY 3

Acid-Base Strength

How is strength measured? In a marathon you need the strength of will to finish. In a Herculean sense, strength is measured by how heavy a load you can lift; in an intellectual sense, by how well you solve problems. A different measuring device is used in each case. In this activity, you will look at how acid strength and base strength are measured.

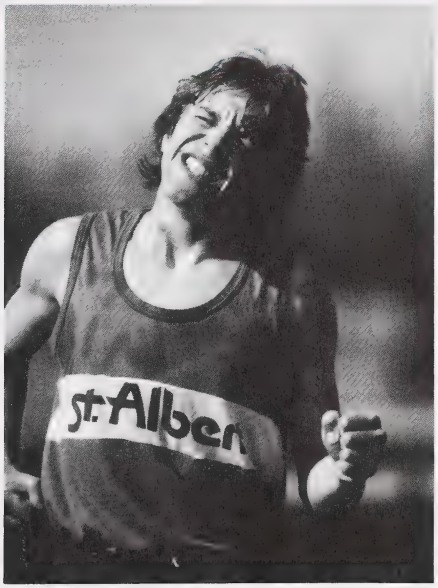

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Table 14.4 on page 449 in your textbook reviews some of the empirical properties of acids and bases. How can those empirical properties help measure the strength of an acid? Does a strong acid turn litmus more red? Does a strong acid react faster with a metal or with carbonates? Lab Exercise 14D: Strengths of Acids investigates one empirical property of acids, pH.

Lab Exercise 14D: Strengths of Acids

Science Skills

- ☑ A. Initiating
- ☐ B. Collecting
- ☐ C. Organizing
- ☑ D. Analysing
- ☑ E. Synthesizing
- ☑ F. Evaluating

Read Lab Exercise 14D on page 449 in your textbook and answer the following questions to complete the exercise. Pay special attention to the applied science skills.

1. Write an experimental design that would generate the evidence in the table. List the controlled, manipulated, and responding variables.

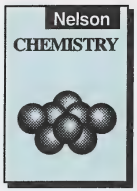
2. Analyse the data by listing the acids from strongest to weakest.

3. Give two other empirical properties listed in Table 14.4 in your textbook that could have been tested in this investigation, and predict the results.

4. Why is it important to keep the temperature and concentration constant when comparing different acids?

Check your answers by turning to the Appendix, Section 2: Activity 3.

From the preceding lab exercise, it is obvious that different acids can show different levels of acidity or strength.



It is important for you to realize there is a difference between the terms *strength* of an acid and *concentration* of an acid. Read the section Strong and Weak Acids on pages 449-450 in your textbook to learn about this difference; then, answer the following questions.

5. Complete the following table comparing the empirical properties of strong and weak acids.

Empirical Properties	Strong Acid (0.10 mol/L)	Weak Acid (0.10 mol/L)
conductivity		
pH		
rate of reaction		

6. Copy the words in the two columns into your notebook; draw lines from the terms in the left-hand column to the appropriate descriptors in the right-hand column.

	<50% reaction with water
Strong acids	100% reaction with water
	products favoured
	reactants favoured
Weak acids	small K
	large K

Percent reaction was the first way you quantitatively measured the extent of an equilibrium (in Section 1, Activity 2 of this module). Weak acid strength is measured in the same way using percent ionization.

$$\text{percent reaction} = \frac{\text{amount of product at equilibrium}}{\text{maximum product possible}} \times 100$$

For Acids,

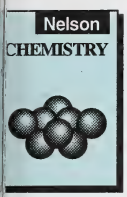
$$\text{percent ionization} = \frac{[H^+_{(aq)}] \text{ at equilibrium}}{\text{original [acid]}} \times 100$$

Notice that if all the acid had ionized, then the maximum $[H^+_{(aq)}]$ possible is the same as the original [acid].

7. Rearrange the following equation to solve for $[H^+_{(aq)}]$ if you were given the percent ionization.

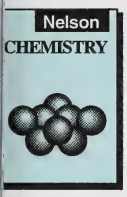
$$\text{percent ionization} = \frac{[H^+_{(aq)}] \text{ at equilibrium}}{\text{original [acid]}} \times 100$$

Check your answers by turning to the Appendix, Section 2: Activity 3.



Read page 451 in your textbook to see sample calculations for finding $[H^+_{(aq)}]$ when given the percent ionization. Try the calculations yourself; then answer the following questions.

8. Calculate the pH of the 0.10 mol/L solutions of $HCl_{(aq)}$ and $CH_3COOH_{(aq)}$ using the $[H^+_{(aq)}]$ calculated in the examples on page 451.



In the Appendix of your textbook (page 553) is a table listing some common acids (by name and formula) from strongest to weakest. The table also includes the percent ionization reaction in 0.10 mol/L solutions, the equilibrium constant for the acid, and the name and formula of the anion produced along with $H^+_{(aq)}$ when the acid ionizes in water.

9. a. How many strong acids are there? List them. (Refer to the Acids and Bases table on page 553 of your textbook.)

   b. Fill in the missing information in the following table. Use the percent ionization listed in the textbook (Appendix, page 553) to calculate the $[H^+_{(aq)}]$ and pH for each of the 0.10 mol/L acids.

Name of Acid	Chemical Formula	Percent Ionization	$[H^+_{(aq)}]$ (mol/L)	pH
hydrocyanic acid				
hydrofluoric acid				
phosphoric acid				
hypochlorous acid				
hydrobromic acid				
nitrous acid				

Your Chemistry Data Booklet (located at the end of the Module 7 Appendix) also has an acid-base table. The data booklet table is different in that it does not have a percent ionization column. You should familiarize yourself with this data booklet table as it is the same as the table you will use when writing the Chemistry 30 diploma examination.

Practically, the pH is measured and used to calculate the percent ionization more often than the other way around. That is, knowing the pH and the concentration of an acid, you can calculate the percent ionization.

10. Make a flowchart that indicates the steps involved in calculating the percent ionization for an acid given the concentration of the acid and the measured pH.

Check your answers by turning to the Appendix, Section 2: Activity 3.

11. Using the steps from your flowchart, calculate the percent ionization for each of the weak acids in the following table.

Acid (name and formula)	Acid Concentration (mol/L)	Measured pH	Percent Ionization
ethanoic, $CH_3COOH_{(aq)}$	0.15	2.78	
ethanoic, $CH_3COOH_{(aq)}$	0.050	3.02	
ethanoic, $CH_3COOH_{(aq)}$	0.010	3.37	
hydrofluoric, $HF_{(aq)}$	0.15	2.00	
hydrofluoric, $HF_{(aq)}$	0.050	2.26	
hydrofluoric, $HF_{(aq)}$	0.010	2.64	

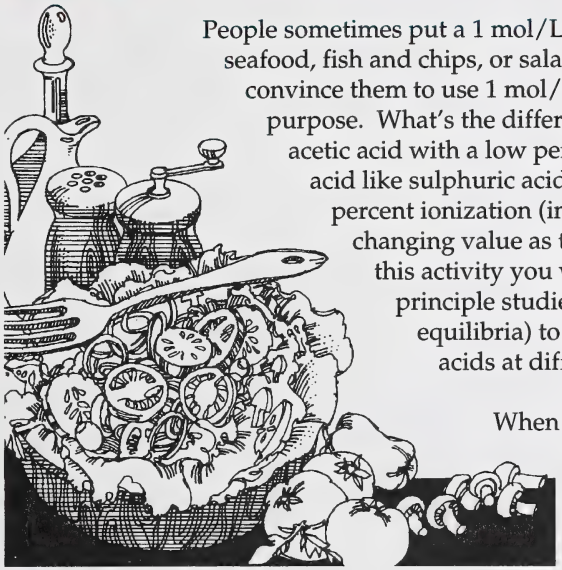
12. Based on the table you have just completed, what generalization can you make about how the percent ionization of a weak acid is related to its concentration?

Check your answers by turning to the Appendix, Section 2: Activity 3.

As you just confirmed, the percent ionization of a weak acid changes depending on the concentration of the acid. This was one of the disadvantages noted earlier with using the percent ionization as a measure of the extent of a reaction. In the final activity, you will learn to use an equilibrium law to establish the extent of the reaction for weak acids in water.

ACTIVITY

4 Equilibrium Constants



People sometimes put a 1 mol/L acetic acid solution (vinegar) on their seafood, fish and chips, or salad, but you would be hard-pressed to convince them to use 1 mol/L sulphuric acid for the same purpose. What's the difference? Vinegar is a solution of weak acetic acid with a low percent ionization compared to a strong acid like sulphuric acid. One disadvantage you found with percent ionization (in Activity 3) was its unpredictable changing value as the acid concentration changed. In this activity you will use the equilibrium constant principle studied in Section 1 (as it applies to acid equilibria) to successfully predict the pH values for acids at different concentrations.

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When a weak acid dissolves in water, it undergoes two processes; it dissolves into molecules and some of the molecules ionize in water. Just how many of the molecules ionize in water depends on the strength of the acid. Read the top half of page 452 in your textbook to learn how the acid ionization constant, K_a, is expressed from the ionization equation for an acid in water; then answer the following questions.

cid ionization onstant (K_a) – quilibrium onstant for the onization of an cid in water

1. Write the ionization equation and the equilibrium law for the following weak acids: $HCN_{(aq)}$, $HCOOH_{(aq)}$, and $H_3PO_{4(aq)}$.

2. What is the advantage of using the K_a value to communicate the strength of an acid (the extent of the equilibrium) over the percent ionization?

Check your answers by turning to the Appendix, Section 2: Activity 4.

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The K_a expression may be more useful in predicting $\left[H^+_{(aq)}\right]$ over a larger range of concentrations, but it can look a bit scary when considering how to solve for $\left[H^+_{(aq)}\right]$. However, there is a way to simplify these calculations. The following example, using the same acid as described on the bottom of page 452 in your textbook and the concepts you studied in Section 1, will show you how this simplification is achieved.

Example

Question: Calculate the $[H^+_{(aq)}]$ and pH of a 1.0 mol/L acetic acid solution.

Answer:

As your textbook suggests, you can't use the percent ionization value from your acid-base table to solve this problem. The percent ionization table value is only suitable for 0.10 mol/L acid solutions. Therefore, you have to use the equilibrium law and the K_a value for acetic acid.

$$CH_3COOH_{(aq)} \rightleftarrows H^+_{(aq)} + CH_3COO^-_{(aq)}$$

$$K_a = \frac{[H^+_{(aq)}][CH_3COO^-_{(aq)}]}{[CH_3COOH_{(aq)}]}$$

Remember, since the K_a expression is an *equilibrium* expression, the concentrations used must be *equilibrium* concentrations at a constant temperature (25°C). How do you get the equilibrium concentrations? You can set up a chart to find the concentrations, just as you did in Section 1.

	$CH_3COOH_{(aq)} \rightleftarrows H^+_{(aq)} + CH_3COO^-_{(aq)}$		
(1) initial concentrations (mol/L)	1.0	___	___
(2) changed concentrations (mol/L)	$-x$	x	x
(3) equilibrium concentrations (mol/L)	$1.0 - x$	x	x

(1) *initial concentrations* – When you first introduce the $CH_3COOH_{(aq)}$ into solution, there are no $H^+_{(aq)}$ or $CH_3COO^-_{(aq)}$.

(2) *changed concentrations* – When the $CH_3COOH_{(aq)}$ ionizes, the following points apply:

- You get a small, but equal, amount of $H^+_{(aq)}$ and $CH_3COO^-_{(aq)}$. The amount of $H^+_{(aq)}$ and the amount of $CH_3COO^-_{(aq)}$ produced is equal to the amount of $CH_3COOH_{(aq)}$ ionized because all mole ratios are 1:1.

- You don't know how much $CH_3COOH_{(aq)}$ is ionized or how much $H^+_{(aq)}$ and $CH_3COO^-_{(aq)}$ is produced. Since this value is unknown, you can set it as ***x***, a variable. Since $CH_3COOH_{(aq)}$ is ionized, let the amount it ionizes equal $-x$; since $H^+_{(aq)}$ and $CH_3COO^-_{(aq)}$ are produced, each of their values is $+x$ (or just x). Eventually, when you solve for x, you will have the required $\left[H^+_{(aq)}\right]$ which can then be converted to pH.

(3) *equilibrium concentrations* – These are the values you can substitute into the equilibrium law expression.

$$K_a = \frac{\left[H^+_{(aq)}\right]\left[CH_3COO^-_{(aq)}\right]}{\left[CH_3COOH_{(aq)}\right]} \quad \rightarrow 1.8\times10^{-5}\ \text{mol/L} = \frac{(x)(x)}{(1.0\ \text{mol/L} - x)}$$

↑ from acid-base table

Normally, this type of mathematical equation would require the use of the quadratic equation to solve for x. But there is a shortcut you can use that eliminates the x from the denominator (bottom) of the equation.

$$1.8\times10^{-5}\ \text{mol/L} = \frac{(x)(x)}{(1.0\ \text{mol/L} - \cancel{x})}$$

↑ cancel

The shortcut is based on the following logic:

- The amount of acid dissociated, x, is relative to the size of the K_a. If the K_a is small, the x value is small.
- If the x value is *small enough*, which occurs when the K_a is at least 1000 times smaller than the acid concentration, it makes very little difference to the acid concentration when x is subtracted.

$$1.8\times10^{-5}\ \text{mol/L} = \frac{(x)(x)}{1.0\ \text{mol/L}\boxed{-x}}$$

↑ **makes little difference**

When calculated by the quadratic equation, the x value is actually equal to 0.004 23. This makes little difference when subtracted from 1.0. Also notice how close this x value is to the x value estimated in the following calculation.

- The mathematical operation is now simplified enough to be solved with less effort.

$$1.8\times10^{-5}\ \text{mol/L} = \frac{(x)(x)}{1.0\ \text{mol/L}} \leftarrow \text{Remember, the } -x \text{ is gone.}$$

$$1.8\times10^{-5}\ \text{mol/L}\ (1.0\ \text{mol/L}) = x^2$$

$$x = \sqrt{1.8\times10^{-5}\ (\text{mol/L})^2}$$

$$= 4.2\times10^{-3}\ \text{mol/L}$$

From your chart, x represents the $\left[H^+_{(aq)}\right]$, which was required.

$$x = \left[H^+_{(aq)}\right] = 4.2\times10^{-3}\ \text{mol/L}$$

$$\text{pH} = -\log\left[H^+_{(aq)}\right]$$

$$= -\log\left(4.2\times10^{-3}\right)$$

$$= 2.37$$

The $\left[H^+_{(aq)}\right]$ is 4.2×10^{-3} mol/L and the pH is 2.37.

Therefore, the general guideline you follow is as follows:

If the [acid] is at least 1000 (one thousand) times greater than the K_a, the $\left[H^+_{(aq)}\right]$ can be solved as shown:

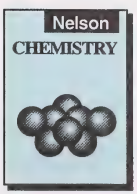
$$K_a = \frac{\left[H^+_{(aq)}\right]^2}{[\text{acid}]} \rightarrow \left[H^+_{(aq)}\right] = \sqrt{K_a \times [\text{acid}]}$$

Note: The preceding discussion and calculations are based on the assumption that one mole of $H^+_{(aq)}$ is produced for every one mole of acid ionized.

3. Under what set of conditions can you make the assumption that the concentration of the acid at equilibrium is the same as the original concentration?

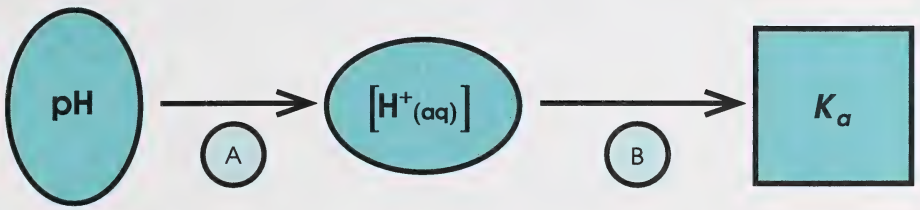
4. Calculate the $\left[H^+_{(aq)}\right]$ and pH of a 0.15 mol/L acetic acid solution.

5. Do questions 32 and 33 on page 458 of your textbook.



Check your answers by turning to the Appendix, Section 2: Activity 4.

In many cases, you will measure the pH of an acid of known concentration, and use it to find the K_a value, rather than doing the reverse operation as shown in the preceding examples. The pH can be used to find the $[H^+_{(aq)}]$ which then gets substituted into the K_a expression along with the original concentration of the acid. The following diagram shows the order in which this is done.



6. For each of the arrows labelled (A) and (B) in the preceding diagram, give the calculation necessary to move along the arrow.

Again, it is important to remember that in the ionization of any acid,

$$HA_{(aq)} \rightleftarrows H^+_{(aq)} + A^-_{(aq)},$$

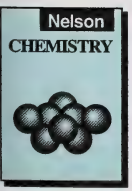
($HA_{(aq)}$: general formula for any acid; $A^-_{(aq)}$: acid ion)

the concentration of the $H^+_{(aq)}$ ions and $A^-_{(aq)}$ ions are *equal*, which simplifies your K_a calculation. Also, the assumption that the [acid] is essentially unchanged simplifies the calculation.

$$HA_{(aq)} \rightleftarrows H^+_{(aq)} + A^-_{(aq)}$$

$$K_a = \frac{[H^+_{(aq)}][A^-_{(aq)}]}{[HA_{(aq)}]} \longrightarrow \text{since } [H^+_{(aq)}] = [A^-_{(aq)}]$$

$$= \frac{[H^+_{(aq)}]^2}{[HA_{(aq)}]}$$



If you know the pH and the $[HA_{(aq)}]$, you can solve for the K_a value. Study the example of this type of calculation (finding K_a) given on the top half of page 453 in your textbook; then do the following questions.

7. Complete the following table.

Acid	[acid] (mol/L)	pH	$[H^+_{(aq)}]$ (mol/L)	K_a (mol/L)
$HCN_{(aq)}$	0.010	5.60		
$HCOOH_{(aq)}$	0.25	2.17		
$CH_3COOH_{(aq)}$	0.15	2.78		
$HClO_{(aq)}$	0.050	4.42		
$HNO_{2(aq)}$	0.12	2.03		

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8. Do question 34 on page 458 in your textbook.

The K_a expression can also be used to find the original concentration of the acid if the K_a value is known and the pH is measured.

9. Rearrange the K_a expression to solve for the [acid].

10. What is the $[HCOOH_{(aq)}]$ in a sample found to have a pH of 2.33? (Remember, the K_a values are listed on page 553 of your textbook.)

11. Do questions 28-31 on page 454 of *Nelson Chemistry*.

Check your answers by turning to the Appendix, Section 2: Activity 4.

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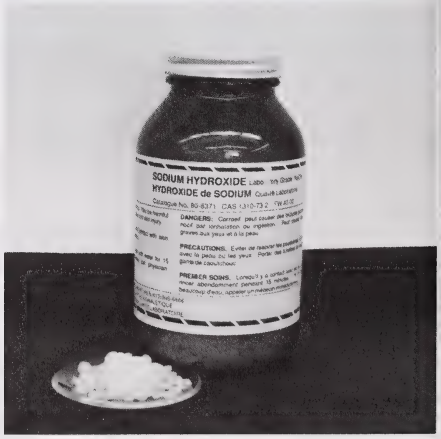
In considering the strength of acids, you need to think as well about the strength of bases. In this activity, you will only consider strong bases (ionic hydroxides). To look at the properties of strong bases, read the bottom half of page 453 in your textbook and answer the following questions.



12. a. Why are all ionic hydroxides considered strong bases?

    b. What property is responsible for the variation in basicity of hydroxides?

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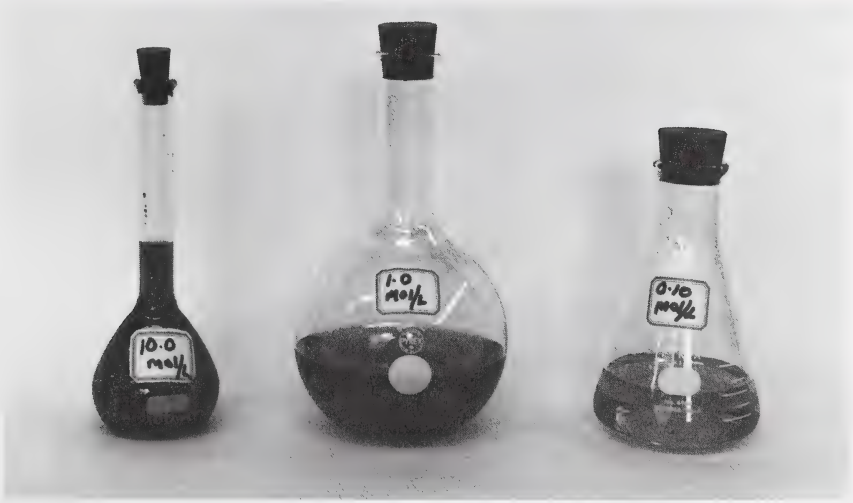
13. Using the dissociation equations for $NaOH_{(aq)}$ and $Ba(OH)_{2(aq)}$ given on page 453 of *Nelson Chemistry*, compare the $[OH^-_{(aq)}]$, pH, and pOH for 0.10 mol/L solutions of each base.

Check your answers by turning to the Appendix, Section 2: Activity 4.



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How could you determine the identity of several unknown solutions? Determining the identity of unknown substances by comparing empirical properties is known as qualitative analysis. Lab Exercise 14E on pages 453-454 in your textbook is an example of using the empirical properties of strong acids, strong bases, and weak acids to perform a qualitative analysis.

Lab Exercise 14E: Qualitative Analysis

Science Skills

- ☐ A. Initiating
- ☐ B. Collecting
- ☐ C. Organizing
- ☑ D. Analysing
- ☐ E. Synthesizing
- ☑ F. Evaluating

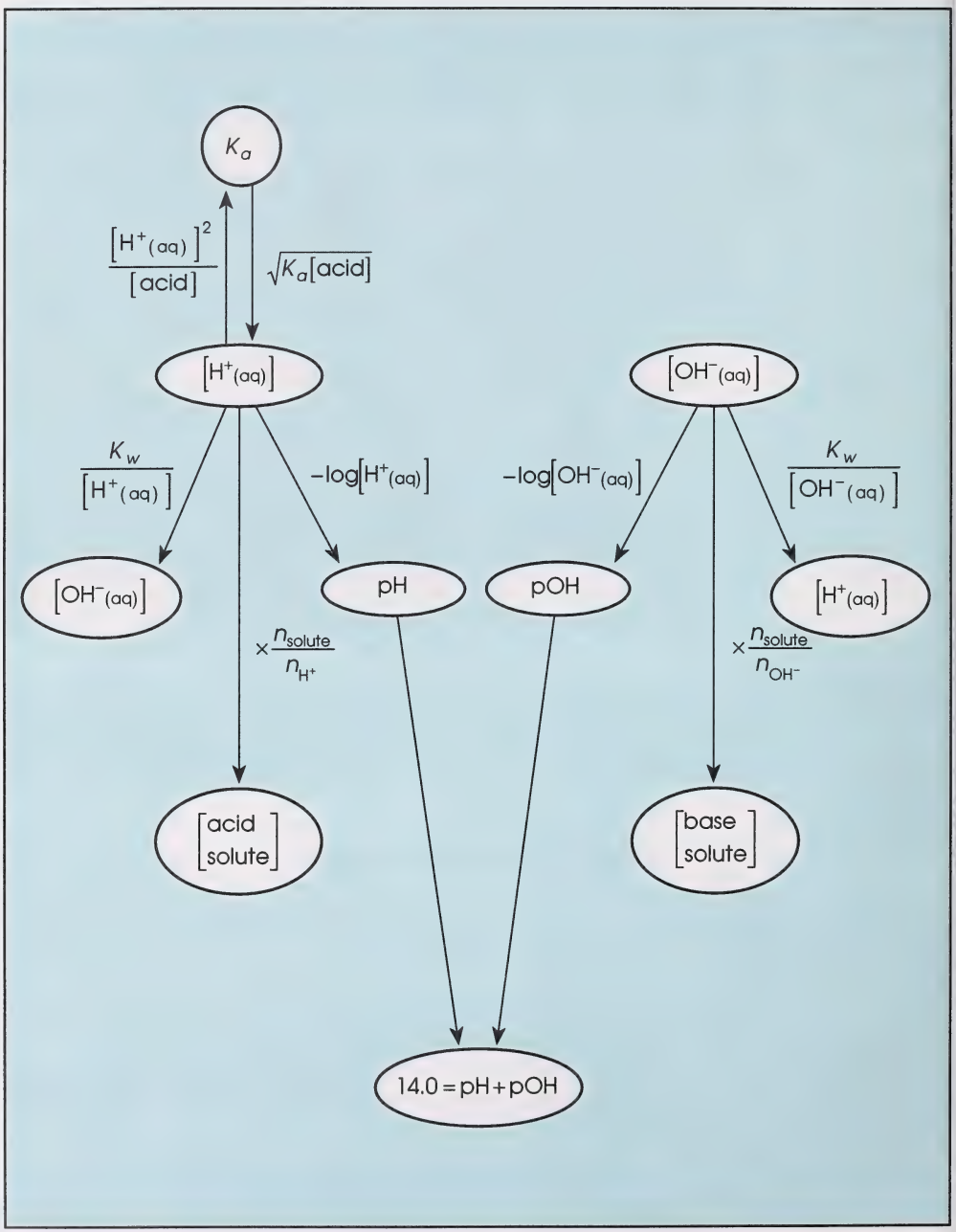
Read the problem, experimental design, and evidence for Lab Exercise 14E. Pay special attention to the required science skills.

14. Complete the analysis for the lab exercise. For each solution, which (if any) was the most important piece of evidence used to classify the solution?

15. Do question 26 (concerning Lab Exercise 14E) on page 454 of your textbook and question 30 on page 458.

Check your answers by turning to the Appendix, Section 2: Activity 4.

In this activity, you have applied the equilibrium law discovered in Section 1, Activity 2 to a specialized solution equilibrium – that of weak acids – and have learned to use the K_a value to predict the $[H^+_{(aq)}]$, pH, and pOH of acids.



Follow-up Activities

If you had difficulties understanding the concepts in the activities, it is recommended that you do the Extra Help. If you have a clear understanding of the concepts, it is recommended that you do the Enrichment.

Extra Help

Complete the following questions as a review of the concepts studied in this section.

1. Match each term in the left-hand column with the best description in the right-hand column.

Terms	Description
a. pH	i. 10^{-pOH}
b. K_w	ii. pOH
c. 14	iii. $[H^+_{(aq)}]$
d. $-\log[OH^-_{(aq)}]$	iv. 0.10 mol/L $NaOH_{(aq)}$
e. 10^{-pH}	v. $[H^+_{(aq)}][OH^-_{(aq)}]$
f. $[OH^-_{(aq)}]$	vi. 0.1 mol/L $HCl_{(aq)}$
g. K_a for $HCN_{(aq)}$	vii. $\frac{[H^+_{(aq)}][CN^-_{(aq)}]}{[HCN_{(aq)}]}$
h. pH = 7	viii. pH + pOH
i. pOH = 1.00	ix. $C_{12}H_{22}O_{11(aq)}$
j. pH = 1.0	x. $-\log[H^+_{(aq)}]$

2. You are given the following question:

> Calculate the mass of $NaOH_{(s)}$ required to make 2.00 L of a solution with a pH of 13.00. Copy the given answer and explain what is occurring beside each step.

a. **Step 1**

$$\begin{aligned} pOH &= 14.00 - pH \\ &= 14.00 - 13.00 \\ &= 1.00 \end{aligned}$$

Explanation: _____

b. **Step 2**

$$\begin{aligned} [OH^-_{(aq)}] &= 10^{-pOH} \\ &= 10^{-1.00} \\ &= 1.0 \times 10^{-1} \text{ mol/L} \end{aligned}$$

Explanation: _____

c. **Step 3**

$$\begin{aligned} n_{(OH^-)} &= \frac{1.0 \times 10^{-1} \text{ mol } OH^-}{\cancel{L}} \times 2.00 \cancel{L} \\ &= 2.0 \times 10^{-1} \text{ mol } OH^- \end{aligned}$$

Explanation: _____

d. **Step 4**

$$NaOH_{(s)} \rightleftarrows Na^+_{(aq)} + OH^-_{(aq)}$$

$$\begin{aligned} n_{(NaOH)} &= 2.0 \times 10^{-1} \cancel{\text{mol } OH^-} \times \frac{1 \text{ mol NaOH}}{1 \cancel{\text{mol } OH^-}} \\ &= 2.0 \times 10^{-1} \text{ mol NaOH} \end{aligned}$$

Explanation: _____

e. **Step 5**

$$\begin{aligned} m_{(NaOH)} &= 2.0 \times 10^{-1} \cancel{\text{mol NaOH}} \times \frac{40.00 \text{ g}}{1 \cancel{\text{mol NaOH}}} \\ &= 8.0 \text{ g} \end{aligned}$$

Explanation: _____

> Keep in mind that Steps 3, 4, and 5 could have been written in one line as follows:
>
> $$m_{(NaOH)} = \frac{1.0 \times 10^{-1} \cancel{\text{mol } OH^-}}{\cancel{L}} \times 2.00 \cancel{L} \times \frac{1 \cancel{\text{mol NaOH}}}{1 \cancel{\text{mol } OH^-}} \times \frac{40.00 \text{ g}}{1 \cancel{\text{mol NaOH}}}$$
> $$= 8.0 \text{ g}$$

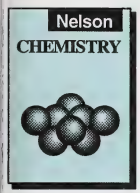
3. If the pH is <7, then the pOH is _____ and the substance is a(n) _____. If the pOH is <7, then the pH is _____ and the substance is a(n) _____.

4. Fill in the right-hand column of the following table using $C_6H_{12}O_{6(aq)}$, $Ba(OH)_{2(aq)}$, $NaCl_{(aq)}$, $HCl_{(aq)}$, and $CH_3COOH_{(aq)}$ as the substances.

Diagnostic Test Results	Substance
turned litmus red, conducted well	
turned litmus blue, conducted well	
no change in litmus, conducted well	
turned litmus red, poor conductor	
no change in litmus, no conductivity	

5. Write out the two methods you discovered in this section by which the $[H^+_{(aq)}]$ can be calculated. Describe any condition either method has to meet.

Check your answers by turning to the Appendix, Section 2: Extra Help.



If you would like further practice in working with the concepts presented in this section, answer any of questions 8, 11-13, 17, 21-29, 31 on pages 456-458 of your textbook. The answers to these questions can be found on pages 518-519 of *Nelson Chemistry*.

Enrichment

Answer **one** of the following questions.

1. When using the K_a expression to find the pH of weak acids, you use an approximation which is valid most of the time. This makes the calculations much simpler, but there are circumstances where the approximation is not true. Instead of assuming the [acid] after ionizing is the same as the [acid] before ionizing, you have to subtract the concentration of ionized acid molecules. This means using the quadratic equation. Use the quadratic equation to find the pH of a 0.20 mol/L solution of sulphurous acid. (Assume that only the first hydrogen ionizes.) The quadratic equation can be found on the page entitled Miscellaneous in your Chemistry Data Booklet (Appendix, Module 7).

2. The solubility of many salts (ionic compounds) are pH dependent in that some dissolve better in higher pH solutions, and some dissolve better in low pH solutions. The following two equations give the reactions responsible for the formation of stalagmites and stalactites in limestone caves.

   Reaction 1: $CO_{2(aq)} + H_2O_{(l)} \rightleftarrows H^+_{(aq)} + HCO_3^-{}_{(aq)}$

   Reaction 2: $H^+_{(aq)} + CaCO_{3(s)} \rightleftarrows Ca^{2+}_{(aq)} + HCO_3^-{}_{(aq)}$

   a. Explain why limestone $\left(CaCO_{3(s)}\right)$ dissolves in water containing $CO_{2(aq)}$, but then reforms in its solid form as stalagmites and stalactites as the water drips from the ceilings of caves (losing $CO_{2(g)}$ to the air).

   b. Many old buildings and statues are made of limestone. The same equilibria as in limestone caves can affect these monuments. Using equilibrium concepts, explain how acidified rain speeds up the deterioration of limestone structures.

   c. During hot weather, farmers sometimes give their chickens carbonated (soda) water to replenish $CO_{2(g)}$ levels in the blood. Since eggshells are made up mainly of $CaCO_{3(s)}$, use the equilibrium given for limestone caves, along with the reaction $CO_{2(g)} \rightleftarrows CO_{2(aq)}$, to explain how farmers are making use of Le Châtelier's principle.

Check your answers by turning to the Appendix, Section 2: Enrichment.

Conclusion

In this section, you have looked at the very important equilibrium that is found in all water systems. You discovered how to measure the $\left[H^+_{(aq)}\right]$ and $\left[OH^-_{(aq)}\right]$ using pH and pOH and you saw how acids and bases affect the water equilibrium and some different ways of expressing their effects. You explored a few examples of how equilibrium principles affect the world around you.

With your new-found knowledge in this module, you should be able to see how equilibrium is important in many other applications.

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Section 2 Assignment: Water's Balance of Power

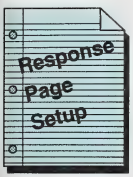
Review the Evaluation information found in the introductory pages of this module.

It is important to number and clearly identify each page with the following information at the top:

Chemistry 30 – Module 5 Section 2 Assignment Page # Name and ID#

Be sure to write legibly. Leave a wide left margin and number all of your pages.

Part A: Multiple Choice



Read each question carefully and decide which of the choices best completes the statement or answers the question. Place the number of each question in a vertical column and place your answer (A, B, C, or D) beside each number. Each question is worth one mark. **(Total = 5 marks)**

1. An Arrhenius acid

   A. is always a neutral molecule
   B. releases hydrogen ions in aqueous solution
   C. is an ionic compound which contains hydrogen ions
   D. produces an aqueous solution whose pH is greater than 7

2. A 12 mol/L solution of acetic acid $\left(CH_3COOH_{(aq)}\right)$ is best described as

   A. a concentrated strong acid
   B. a concentrated weak acid
   C. a dilute strong acid
   D. a dilute weak acid

3. Hydroiodic acid is a very strong acid. In an aqueous solution of 0.015 mol/L $HI_{(aq)}$, the $\left[H^+_{(aq)}\right]$ is approximately

   A. 10^{-13} mol/L
   B. 10^{12} mol/L
   C. 10^{-2} mol/L
   D. 10^{-1} mol/L

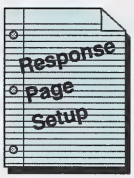
4. What is the $[OH^-_{(aq)}]$ in a 0.20 mol/L $HNO_{3(aq)}$ (a strong acid) solution?

   A. 2×10^{-1} mol/L
   B. 5×10^{-7} mol/L
   C. 2×10^{-14} mol/L
   D. 5×10^{-14} mol/L

5. A solution has a pH of 5. The $[H^+_{(aq)}]$ and $[OH^-_{(aq)}]$ in the solution respectively are

   A. 10^{-5} and 10^{-9} mol/L
   B. 10^{-9} and 10^{-5} mol/L
   C. 10^{+5} and 10^{+9} mol/L
   D. 10^{+9} and 10^{+5} mol/L

Part B: Written Response



Write your answers neatly and concisely. For full marks, your answers must show all pertinent explanations, calculations, and formulas.

6. a. Write the ionization equation for benzoic acid $C_6H_5COOH_{(aq)}$. **(1 mark)**

   b. Calculate the pH of a 0.0150 mol/L solution of $C_6H_5COOH_{(aq)}$ using the K_a expression. Show your work. **(3 marks)**

7. Household vinegar is labelled as being 5.0% by mass (in every 100 mL of vinegar solution, there should be 5.0 g of acetic acid, $CH_3COOH_{(aq)}$). To test this claim, Lane measured the pH of a sample of household vinegar and found it to be 2.435.

   a. Analyse Lane's results by calculating the concentration of acetic acid in vinegar in moles per litre. Show your work. **(3 marks)**

   b. Find the actual mass of acetic acid in 100 mL of Lane's vinegar sample. Show your work. **(3 marks)**

   c. Calculate the percent error from the labelled concentration (5% or 5.0 g/100mL). Show your work. **(1 mark)**

8. In an experiment to determine the relative strengths of acids, 8.0 g of $HF_{(g)}$ was dissolved in 2.00 L of water and the resulting solution had a pH of 2.080. Calculate the experimental percent ionization of the hydrofluoric acid. Show your work. **(3 marks)**

9. Methanoic acid, $HCOOH_{(aq)}$, is sometimes known as formic acid (from the Latin *formica* for *ant*, since ants secrete formic acid). Suppose that an ant uses 1.0×10^{-5} mL of formic acid with a pH of 2.456 every time it bites. What mass of methanoic acid does it use in each bite? Show your work. **(4 marks)**

10. Construct a table similar to the one that follows and fill in the blanks.

$[H^+_{(aq)}]$	$[OH^-_{(aq)}]$	pH	pOH	Acid, Base, or Neutral
2.5×10^{-5}				
	8×10^{-11}			
		11.3		
			3.46	
				neutral

(10 marks)

Science Skills
- ☑ A. Initiating
- ☐ B. Collecting
- ☐ C. Organizing
- ☐ D. Analysing
- ☐ E. Synthesizing
- ☐ F. Evaluating

11. Your solubility table will tell you that $CH_3COO^-_{(aq)}$ ions are insoluble with $Ag^+_{(aq)}$ ions (i.e., they will precipitate out of solution).

   a. Make a prediction about the effect of $Ag^+_{(aq)}$ ions on the pH of acetic acid when $AgNO_{3(aq)}$ is added to $CH_3COOH_{(aq)}$. Give reasons for your prediction. **(3 marks)**

   b. Predict what the pH of the acetic acid solution should be if its concentration is 8.3×10^{-3} mol/L. Show your work. **(3 marks)**

   c. Write an experimental design for an investigation to test what effect the addition of $AgNO_{3(aq)}$ has on the pH of an acetic acid solution. **(3 marks)**

   d. Write a procedure for the investigation. **(3 marks)**

Response Page Setup

Reproduce the science skills assessment box after your response to this question. Remember to indicate your evaluation of your skill level for each identified science skill. Leave a space for the teacher assessment.

Self:	A.	☐	B.	☐	C.	☐	D.	☐	E.	☐	F.	☐
Teacher:	A.	☐	B.	☐	C.	☐	D.	☐	E.	☐	F.	☐

MODULE SUMMARY

WESTFILE INC.

This module has explored reaction systems which are reversible – they are in balance or equilibrium. By learning how to measure and communicate the extent to which a reaction is reversible and how to manipulate chemical equilibria, you have seen how the balance of a system can be influenced to help with large-scale chemical processes, making them economically viable. Finally, applying these concepts to water-based systems has clarified your understanding of what is meant when solutions or products are advertised as being pH balanced.

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Final Module Assignment

Review the Evaluation information found in the introductory pages of this module.

It is important to number and clearly identify each page with the following information at the top:

Chemistry 30 – Module 5 Final Module Assignment Page # Name and ID#

Be sure to write legibly. Leave a wide left margin and number all of your pages.

1. Phosgene gas, $COCl_{2(g)}$, decomposes to $CO_{(g)}$ and $Cl_{2(g)}$ at high temperatures, as shown by the following equilibrium equation:

$$COCl_{2(g)} \rightleftarrows CO_{(g)} + Cl_{2(g)}$$

   a. Write the equilibrium law for the decomposition of phosgene. **(1 mark)**

   b. Using the following table of equilibrium conditions at 400°C, calculate the K value for the system and calculate the missing equilibrium concentrations. Show your work. **(3 marks)**

Trial	$[COCl_{2(g)}]$ (mol/L)	$[CO_{(g)}]$ (mol/L)	$[Cl_{2(g)}]$ (mol/L)
i.	0.0440	0.0060	0.0060
ii.	0.750	0.0040	?
iii.	?	0.025	0.000 15

   c. Calculate the ΔH_r for the decomposition of one mole of phosgene gas into carbon monoxide and chlorine gas. (The $\Delta H°_f$ of $COCl_{2(g)}$ is -218.8 kJ/mol). **(2 marks)**

   d. If the temperature of the phosgene reaction mixture at equilibrium was increased, what effect would it have on the equilibrium constant, K? Would K get larger or smaller? Explain. **(1 mark)**

   e. What direction, if any, would you expect this equilibrium to shift if a sample had its volume suddenly decreased? Explain. **(1 mark)**

   f. A small amount of $CO_{(g)}$ in the equilibrium mixture was replaced by an equal amount of $CO_{(g)}$ labelled with radioactive ^{14}C. If a sample of phosgene gas, $COCl_{2(g)}$, was isolated from the mixture several hours later, would you expect it to contain any of the radioactive carbon? Explain your answer in terms of equilibrium concepts. **(2 marks)**

Science Skills

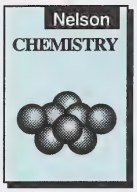
- ☐ A. Initiating
- ☐ B. Collecting
- ☐ C. Organizing
- ☑ D. Analysing
- ☐ E. Synthesizing
- ☐ F. Evaluating

2. Chris has been assigned the task of determining the concentration of a weak acid and then identifying it. Chris knows that the K_a of an acid can be found if the pH and concentration of the acid are known, and has been told by the lab instructor that the acid reacts in a 1:1 mole ratio with NaOH. Chris measured the pH of the acid and then titrated a 25.00 mL sample of the acid with 0.150 mol/L $NaOH_{(aq)}$. The evidence obtained follows in a table.

pH of acid	2.156
Volume of acid titrated	25.00 mL
Concentration of $NaOH_{(aq)}$	0.150 mol/L
Average volume of $NaOH_{(aq)}$ required to completely react with the acid	45.14 mL

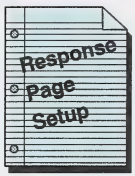
a. What is the $[H^+_{(aq)}]$ of the unknown weak acid? **(1 mark)**

b. Calculate the concentration of the acid. **(2 marks)**



c. Calculate the K_a of the acid. **(1 mark)**

d. Compare the K_a from part c. to the value on the Acids and Bases table on page 553 of your textbook. Determine which acid the unknown could be. **(1 mark)**



Reproduce the science skills assessment box after your response to this question. Remember to indicate your evaluation of your skill level for each identified science skill. Leave a space for the teacher assessment.

Self:	A.	☐	B.	☐	C.	☐	D.	☐	E.	☐	F.	☐
Teacher:	A.	☐	B.	☐	C.	☐	D.	☐	E.	☐	F.	☐

3. Investigation: Solution Equilibria

Science Skills

- ☑ A. Initiating
- ☑ B. Collecting
- ☑ C. Organizing
- ☑ D. Analysing
- ☐ E. Synthesizing
- ☑ F. Evaluating

You will perform an investigation to test various solution equilibria related to a saturated solution of calcium hydroxide (limewater, $Ca(OH)_{2(aq)}$). You will complete the prediction, evidence, analysis, and evaluation in your investigation report. Pay special attention to the required components, safety aspects, and applied science skills.

Problem

What will be the effects of first adding $NaOH_{(aq)}$, and then $HCl_{(aq)}$, to a sample of saturated $Ca(OH)_{2(aq)}$ (limewater) which is at equilibrium?

a. Make a prediction as to what should happen to the limewater equilibrium with the addition of each of the $NaOH_{(aq)}$ and $HCl_{(aq)}$ solutions. Use equations to support your prediction. **(3 marks)**

Guidelines

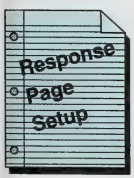
- All the chemicals used in this investigation are corrosive – handle them with care.
- Follow the guidelines for chemical disposal as given earlier in this module.

Materials

- approximately 5 mL of each of the following solutions
  - saturated $Ca(OH)_{2(aq)}$ (limewater)
  - 0.10 mol/L $NaOH_{(aq)}$
  - 0.10 mol/L $HCl_{(aq)}$
- (2) 13 mm × 100 mm test tubes
- 2 eyedroppers
- rubber stopper for test tubes

Procedure

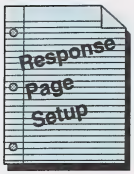
- Add 10 drops of saturated $Ca(OH)_{2(aq)}$ (limewater) to a clean, dry test tube. (You made some limewater in an earlier lab, remember?)
- Slowly add 10 drops of 0.10 mol/L $NaOH_{(aq)}$ to the limewater, observing as you add each drop.
- Stopper the test tube. Swirl the test tube contents.
- Add enough drops of 0.10 mol/L $HCl_{(aq)}$ (again slowly) until a definite change is observed. Stopper the test tube and swirl after each drop.
- Repeat the procedure if times permits.



b. Perform the procedure and record your evidence in a suitable table. **(3 marks)**

c. Analyse your results. **(2 marks)**

d. Evaluate the procedure and Le Châtelier's principle on the basis of your results. **(2 marks)**



Reproduce the science skills assessment box after your response to this question. Remember to indicate your evaluation of your skill level for each identified science skill. Leave a space for the teacher assessment.

Self:	A.	☐	B.	☐	C.	☐	D.	☐	E.	☐	F. ☐
Teacher:	A.	☐	B.	☐	C.	☐	D.	☐	E.	☐	F. ☐

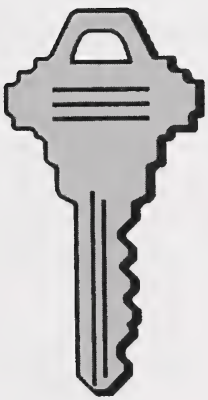
To ensure that all of your work has been completed in a satisfactory manner, check off the items in the following list:

- ❑ Section 1 Assignment has been completed.
- ❑ Section 2 Assignment has been completed.
- ❑ Final Module Assignment has been completed.
- ❑ Your responses are organized and neat, with room for teacher comments.
- ❑ All your response pages are numbered consecutively and identified with this heading.

Chemistry 30 – Module 5 Section # Assignment Page # Name and ID #

Submit **only** your **assignment response pages** for evaluation.

Appendix



Glossary
Suggested Answers

Glossary

acid ionization constant (K_a): equilibrium constant for the ionization of an acid in water

auto-ionization: ionization without external influences

buffer: a solution that maintains a near-constant, specified pH

desiccator: something which dries (removes water from) substances

dynamic equilibrium: a balance between forward and reverse processes occurring at the same rate (also see *equilibrium*)

equilibrium: the condition of a closed system in which all observable properties remain constant at a constant temperature

equilibrium constant (*K*): the value obtained from the mathematical combination of equilibrium concentrations using the equilibrium law

equilibrium law: an expression equal to the product of the equilibrium concentrations of the chemical products divided by the product of the equilibrium concentrations of the chemical reactants, where each concentration is raised to the power of the coefficient of that entity

ionization: the process of converting an atom or molecule to an ion

K_w: equilibrium constant for the ionization of water

macroscopic: that which is observable

percent reaction: the equilibrium concentration of product in an equilibrium as compared to the maximum possible concentration, expressed as a percentage

Suggested Answers

Section 1: Activity 1

1. A typical set of observations might be as follows:

Volume of Vinegar Added (mL)	Temperature of Solution (°C)	Observations
0	22.5	• Baking soda is a white solid. • The baking soda solution is clear and colourless. • The solution turns light blue when HBb* is added.
10	22.6	• There is vigorous bubbling when vinegar is added; 2 cm of foam is on the surface. • The solution turns yellow initially, then blue-green; the bubbles subside.
20	22.8	• There is vigorous bubbling when vinegar is added; 2 cm of foam is on the surface. • The solution turns yellow, then green; bubbles subside.

30	22.9	• There is bubbling when vinegar is added; 1 cm of foam is on the surface. • The solution turns yellow, then yellow-green; bubbles subside.
40	23.0	• There is slight bubbling. • The solution turns bright yellow. • Some bubbles remain.
50	23.0	• There is no apparent change in the solution.

*HBb is the abbreviation for the indicator bromothymol blue.

2. a. The colour of the solution was blue.

   b. Based on what you learned in Chemistry 20, this colour should indicate the presence of a **base**.

3. Yes. The solution exhibits the following constant macroscopic properties:

   - no colour change
   - no temperature change
   - no gas production
   - no composition change

4. The empirical properties that can be measured for any system at equilibrium include:

   - temperature
   - pressure
   - colour
   - composition
   - amounts of substances

5. No, the system was not at equilibrium because the solution was changing colour, the temperature was increasing, and a gas was being produced (change in composition).

6. The gas produced was carbon dioxide. Carbon dioxide is the gas produced when opening a soft drink.

7. In this investigation, the setup is not a closed system with respect to gases because it is not closed on all sides. A soft drink bottle is a closed system for gases. Once the bottle is opened, the system is no longer closed and the gas can escape.

8. The end of the reaction was indicated when no more gas bubbles were forming and when the temperature stopped increasing. (You cannot use colour change because the colour stops changing before the gas stops escaping.)

9. The four assumptions are that all reactions are fast, spontaneous, quantitative, and stoichiometric.

10. Some redox reactions were found to be non-spontaneous.

11. Answers may vary. A suitable procedure might be as follows:

    - Measure 5 mL of 0.10 mol/L of $CaCl_{2(aq)}$ using a clean graduated cylinder and pour the solution into a clean 100 mL beaker.

    - Measure 5 mL of 0.10 mol/L $Na_2SO_{4(aq)}$, using a clean graduated cylinder.

- Slowly add the $Na_2SO_{4(aq)}$ into the beaker containing the $CaCl_{2(aq)}$. Stir constantly.
- Filter the precipitate from the mixture.
- Collect two separate 1 mL samples of the filtrate in test tubes (about 1 cm-2 cm in the bottom of each test tube).
- Test one filtrate sample with a few drops of $Na_2CO_{3(aq)}$ and record the result.
- Test the other filtrate sample with a few drops of $Ba(NO_3)_{2(aq)}$ and record the results.
- Repeat the experiment with varying amounts of $CaCl_{2(aq)}$ and $Na_2SO_{4(aq)}$ such as 5 mL/10mL or 10 mL/5 mL.

12. For the procedure outlined, an evidence table might appear as follows:

Trial #	$CaCl_{2(aq)}$ (mL)	$Na_2SO_{4(aq)}$ (mL)	Test with Na_2CO_3	Test with $Ba(NO_3)_2$
1	5	5	✔	✔
2	10	5	✔	✔
3	5	10	✔	✔

- A white precipitate formed in each trial.
- A white precipitate formed in each of the six diagnostic tests (as indicated by the ✔).

13. The chemical equation for the reaction is as follows:

$$CaCl_{2(aq)} + Na_2SO_{4(aq)} \rightarrow CaSO_{4(s)} + 2\,NaCl_{(aq)}$$

The precipitate would be calcium sulphate, $CaSO_{4(s)}$.

14. For the procedure outlined, the evidence table would appear as follows:

Trial #	Moles of $CaCl_{2(aq)}$	Moles of $Na_2SO_{4(aq)}$	Limiting Reagent	Excess Reagent
1	5×10^{-4}	5×10^{-4}	none	none
2	1×10^{-3}	5×10^{-4}	Na_2SO_4	$CaCl_2$
3	5×10^{-4}	1×10^{-3}	$CaCl_2$	Na_2SO_4

A sample calculation for Trial 2 would be as follows:

$$CaCl_{2(aq)} + Na_2SO_{4(aq)} \rightarrow CaSO_{4(s)} + 2\,NaCl_{(aq)}$$

10 mL 5 mL
0.10 mol/L 0.10 mol/L

$$n_{(CaCl_2)} = 0.010\ \text{L} \times \frac{0.10\ \text{mol}}{\text{L}} = 1 \times 10^{-3}\ \text{mol}$$

$$n_{(Na_2SO_4)} = 0.005\ \text{L} \times \frac{0.10\ \text{mol}}{\text{L}} = 5 \times 10^{-4}\ \text{mol}$$

$CaCl_{2(aq)}$ and $Na_2SO_{4(aq)}$ react in a 1:1 mole ratio (from the equation), but since there is more $CaCl_{2(aq)}$ present (in moles) than $Na_2SO_{4(aq)}$, there should be some $CaCl_{2(aq)}$ left over (excess reagent) and all the $Na_2SO_{4(aq)}$ should be used up (limiting reagent).

15. The positive diagnostic tests show for each trial that both reactants were found in the filtrate. This suggests that neither reactant was completely used up.

16. The experiment was able to prove that this reaction does not react quantitatively. Varying the quantities of reactants helped to confirm this result since a different reactant was the limiting reagent in Trials 2 and 3. Using both diagnostic tests was necessary to confirm that both reactants remained in the filtrate. You can say that not all reactions react quantitatively.

17. To check your results, you would need to perform the experiment more than once.

18.
a. reverse
b. balance
c. equilibrium
d. closed
e. phase
f. solubility
g. $I_{2(aq)}$
h. rate
i. $I_{2(s)}$
j. dynamic
k. liquid
l. rate
m. vapour
n. chemical

19. If it is easier to carry out the forward reaction, it is easier to form products than reactants. You would expect, then, to have more products than reactants at equilibrium.

Section 1: Activity 2

1. When a sealed container of colourless dinitrogen tetraoxide was heated, the gas was converted to brown nitrogen dioxide. When a sealed nitrogen dioxide container was cooled, the brown colour disappeared. When the two containers were left at the same temperature, both containers appeared to be the same colour.

2. a. When hydrogen gas and iodine gas are mixed and then heated, they react to form **hydrogen iodide**. Evidence to indicate this reaction occurs is the **disappearance** of the purple colour associated with **iodine**.

   b. The faint pink colour indicates that some iodine has not been reacted.

   c. The constant colour suggests that equilibrium has been reached.

3. a. In the kinetic molecular theory, molecules are assumed to have these types of motion: rotational, vibrational, and translational.

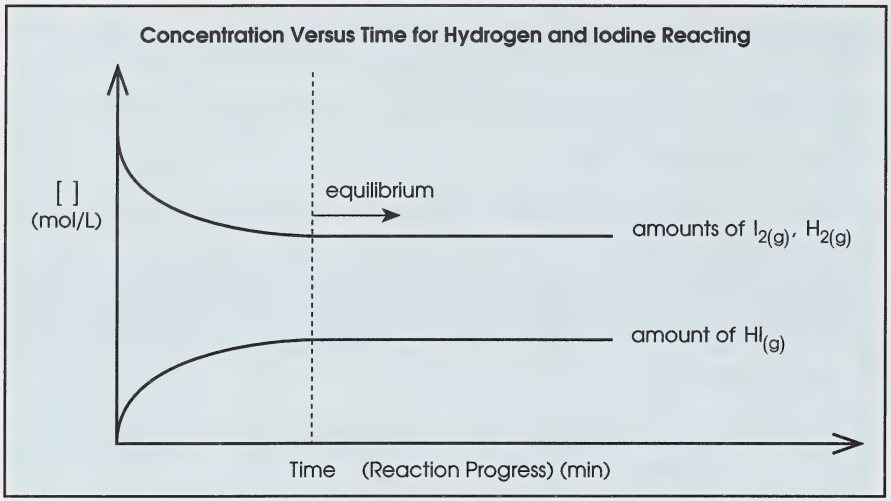
   b. In order for two molecules to react they must meet with sufficient energy and at the correct angle.

4. a. The hydrogen iodide equilibrium develops as follows.

   Initially only the forward reaction occurs but as more hydrogen iodide is produced the reverse reaction starts. As the concentration of reactants decreases, the forward rate decreases; as more products are formed, the reverse rate increases.

   b. When the forward and reverse reaction rates are equal, equilibrium has been reached.

   c. This kind of equilibrium is called a dynamic equilibrium.

5. 

Note: [ ] – means "concentration of"

6. A dynamic equilibrium is indicated by double arrows ($\rightleftarrows$) in a reaction equation.

7. A radioactive isotope of iodine is used in place of normal iodine in a reaction mixture. After a period of time some hydrogen iodide was shown to contain the radioactive iodine. This is evidence that an equilibrium is dynamic.

8. a. The first system has only $H_{2(g)}$ and $I_{2(g)}$ initially present (reactants in equal amounts) and no $HI_{(g)}$ (product). The second system has both reactants (in equal concentrations) and product present. The third system initially has only $HI_{(g)}$ (product) present.

   b. Each system has differing concentrations from the other systems, but all three systems have both reactants and product present at equilibrium.

9. $$\text{percent reaction} = \frac{\boxed{\textbf{equilibrium [product]}}}{\boxed{\textbf{maximum [product]}}} \times 100\%$$

10. THE HYDROGEN-IODIDE SYSTEM AT 448°C

System	Initial $[H_{2(g)}]$	Initial $[I_{2(g)}]$	Initial $[HI_{(g)}]$	Maximum $[HI_{(g)}]$	Equilibrium $[HI_{(g)}]$	Percent Reaction (%)
a	2.0	2.0	1.0	5.0	3.9	78
b	10.0	10.0	0.0	**20.0**	15.6	**78.0**
c	7.50	7.50	2.00	**17.0**	13.3	**78.2**
d	0.40	0.40	1.60	**2.40**	1.87	**77.9**
e	0.00	0.00	1.50	**1.50**	1.18	**78.7**
f	1.20	1.20	1.00	**3.40**	2.65	**77.9**

11. **Textbook question 7:**

a. $3\,H_{2(g)} + N_{2(g)} \overset{11\%}{\rightleftarrows} 2\,NH_{3(g)}$

b. $C_{(s)} + H_2O_{(g)} \overset{>50\%}{\rightleftarrows} CO_{(g)} + H_{2(g)}$

c. $2\,Ag^{+}_{(aq)} + Cu_{(s)} \overset{>99\%}{\rightleftarrows} 2\,Ag_{(s)} + Cu^{2+}_{(aq)}$

d. $2\,SO_{2(g)} + O_{2(g)} \overset{65\%}{\rightleftarrows} 2\,SO_{3(g)}$

Textbook question 8:

$Ni_{(s)} + 2\,Cr^{3+}_{(aq)} \overset{<50\%}{\rightleftarrows} Ni^{2+}_{(aq)} + 2\,Cr^{2+}_{(aq)}$

Less than 50% reaction would correspond to a non-spontaneous reaction.

12. a. The reactants are favoured.

b. The products are favoured.

c. The reaction is quantitative (products are heavily favoured).

d. The products are favoured.

13. THE IRON(III)-THIOCYANATE REACTION SYSTEM

Calculation	Trial 1	Trial 2	Trial 3	Trial 4	Trial 5
$[Fe^{3+}][SCN^{-}][FeSCN^{2+}]$	2.89×10^{-9}	2.34×10^{-9}	1.51×10^{-9}	4.11×10^{-10}	3.52×10^{-10}
$[Fe^{3+}]+[SCN^{-}]+[FeSCN^{2+}]$	4.01×10^{-2}	1.58×10^{-2}	7.29×10^{-3}	3.04×10^{-3}	2.72×10^{-3}
$\frac{[Fe^{3+}][SCN^{-}]}{[FeSCN^{2+}]}$	3.40×10^{-3}	3.41×10^{-3}	3.48×10^{-3}	3.26×10^{-3}	3.38×10^{-3}
$\frac{[FeSCN^{2+}]}{[Fe^{3+}][SCN^{-}]}$	294	293	288	307	296

14. The relationships $\frac{[FeSCN^{2+}]}{[Fe^{3+}][SCN^-]}$ and $\frac{[Fe^{3+}][SCN^-]}{[FeSCN^{2+}]}$ give near constant values.

15. The amounts of reactants and products in equilibrium law expressions are measured in **moles per litre (mol/L)**.

16. When writing the equilibrium law, the **products** are always in the numerator and the **reactants** are always in the denominator.

17. $K = \frac{S^{y}\,T^{z}}{Q^{w}\,R^{x}}$

18. A large K value means a **large** percent reaction and that **products** are favoured. A small K value means a **small** percent reaction and that **reactants** are favoured.

19.

K Value	Classification
5.6	products favoured
0.96	balanced
1.0×10^{12}	quantitative
3.1×10^{-8}	reactants favoured

Note: These K_a ranges and classifications are general in nature only.

20. $K = \frac{[HI_{(g)}]^2}{[H_{2(g)}][I_{2(g)}]}$

21. a. System 1

$$K = \frac{(1.56\ \text{mmol/L})^2}{(0.22\ \text{mmol/L})^2}$$
$$= 50$$

System 2

$$K = \frac{(2.10\ \text{mmol/L})^2}{(0.30\ \text{mmol/L})^2}$$
$$= 49$$

System 3

$$K = \frac{(2.50\ \text{mmol/L})^2}{(0.35\ \text{mmol/L})^2}$$
$$= 51$$

b. $K_{avg} = \frac{50 + 49 + 51}{3}$
$= 50$

22.

EQUILIBRIUM CONDITIONS FOR HYDROGEN-IODINE SYSTEM AT 448°C

System	$[H_{2(g)}]$ (mol/L)	$[I_{2(g)}]$ (mol/L)	$[HI_{(g)}]$ (mol/L)
a.	0.40	0.050	1.00
b.	0.16	9.00	8.5
c.	**0.845**	0.333	3.750
d.	3.56	**2.82**	22.4
e.	4.0*x*	9.0*x*	**42x**

Note: The use of x in System e. is simply to show that any equilibrium concentrations will follow the equilibrium law at constant temperature.

The calculations for Systems c., d., and e. are as follows:

c. $$[H_{2(g)}] = \frac{[HI_{(g)}]^2}{K[I_{2(g)}]}$$
$$= \frac{(3.750 \text{ mol/L})^2}{50 \times 0.333 \text{ mol/L}}$$
$$= 0.845 \text{ mol/L}$$

d. $$[I_{2(g)}] = \frac{[HI_{(g)}]^2}{[H_{2(g)}]K}$$
$$= \frac{(22.4 \text{ mol/L})^2}{(3.56 \text{ mol/L})(50)}$$
$$= 2.82 \text{ mol/L}$$

e. $$[HI_{(g)}] = \sqrt{K[H_{2(g)}][I_{2(g)}]}$$
$$= \sqrt{(50)(4.0x)(9.0x)}$$
$$= \sqrt{1800x^2}$$
$$= 42x$$

23. **Step 1** Write the equilibrium law expression.

$$K = \frac{[CO_{2(g)}][H_{2(g)}]}{[CO_{(g)}][H_2O_{(g)}]}$$

Step 2 Substitute the values into the expression and solve for K.

$$K = \frac{(0.100 \text{ mol/L})(0.300 \text{ mol/L})}{(0.296 \text{ mol/L})(0.0249 \text{ mol/L})}$$
$$= 4.07$$

The equilibrium constant is 4.07.

24. **Step 1** Write the equilibrium law expression.

$$K = \frac{\left[NO_{2(g)}\right]\left[SO_{2(g)}\right]}{\left[SO_{3(g)}\right]\left[NO_{(g)}\right]}$$

Step 2 Rearrange the expression, substitute in the values, and solve for $\left[NO_{(g)}\right]$.

$$\left[NO_{(g)}\right] = \frac{\left[NO_{2(g)}\right]\left[SO_{2(g)}\right]}{\left[SO_{3(g)}\right]K}$$

$$= \frac{(0.15 \text{ mol/L})(0.40 \text{ mol/L})}{(0.48 \text{ mol/L})(0.50)}$$

$$= 0.25 \text{ mol/L}$$

The equilibrium $\left[NO_{(g)}\right]$ is 0.25 mol/L.

25. **Step 1** Write the equilibrium law expression.

$$K = \frac{\left[H_{2(g)}\right]\left[Br_{2(g)}\right]}{\left[HBr_{(g)}\right]^2}$$

Step 2 Calculate the initial $\left[HBr_{(g)}\right]$.

$$\text{initial } [HBr] = \frac{0.750 \text{ mol}}{1.50 \text{ L}}$$

$$= 0.500 \text{ mol/L}$$

Step 3 Set up an information table to organize your concentration values.

	$2\,HBr_{(g)}$	$\rightleftarrows$	$H_{2(g)}$	+	$Br_{2(g)}$
initial concentrations (mol/L)	0.500		0.000		0.000
changed concentrations (mol/L)	−2(0.130)		0.130		0.130
equilibrium concentrations (mol/L)	0.240		0.130		0.130

Step 4 Substitute the equilibrium concentrations into the equilibrium law expression and solve for *K*.

$$K = \frac{(0.130 \text{ mol/L})(0.130 \text{ mol/L})}{(0.240 \text{ mol/L})^2}$$
$$= 0.293$$

The equilibrium constant is 0.293.

26.

Method	Advantages	Disadvantages
Percent reaction	• easy to understand • easy to calculate	• cannot be used to calculate for new conditions • changes with concentration
Equilibrium constant	• more flexible – can be used to predict new conditions	• depends on temperature • does not describe rate of reaction • affected by large changes in concentration

27. **Textbook question 9:**

The equilibrium law only describes what the equilibrium conditions are; it does not explain why they are what they are. It gives a calculable quantity associated with the equilibrium.

Textbook question 10.:

a. $H_{2(g)} + Cl_{2(g)} \rightleftarrows 2\,HCl_{(g)}$ $\quad K = \dfrac{\left[HCl_{(g)}\right]^2}{\left[H_{2(g)}\right]\left[Cl_{2(g)}\right]}$

b. $N_{2(g)} + 3\,H_{2(g)} \rightleftarrows 2\,NH_{3(g)}$ $\quad K = \dfrac{\left[NH_{3(g)}\right]^2}{\left[H_{2(g)}\right]^3\left[N_{2(g)}\right]}$

c. $2\,H_{2(g)} + O_{2(g)} \rightleftarrows 2\,H_2O_{(g)}$ $\quad K = \dfrac{\left[H_2O_{(g)}\right]^2}{\left[H_{2(g)}\right]^2\left[O_{2(g)}\right]}$

d. $Ni^{2+}{}_{(aq)} + 6\,NH_{3(aq)} \rightleftarrows Ni(NH_3)_6{}^{2+}{}_{(aq)}$ $\quad K = \dfrac{\left[Ni(NH_3)_6{}^{2+}{}_{(aq)}\right]}{\left[Ni^{2+}{}_{(aq)}\right]\left[NH_{3(aq)}\right]^6}$

Section 1: Activity 3

1. Manufacturers increase the pressure in the soft drink container, which forces more carbon dioxide to dissolve in solution.

2. Le Châtelier's principle states that if a system at equilibrium is subjected to stress, the system acts to relieve the effects of the stress..

3. $N_{2(g)} + 3\,H_{2(g)} \rightleftarrows 2\,NH_{3(g)} + 92\text{ kJ}$

4. The formation of ammonia from hydrogen and nitrogen is an **exothermic** reaction; the decomposition of ammonia is an **endothermic** reaction.

5. When heat is added to a closed container of ammonia, nitrogen, and hydrogen at equilibrium, the equilibrium will shift to **use up** energy and relieve the stress. Since the **decomposition** of ammonia uses up energy, the **reactants** are favoured by the addition of heat.

6. No. The addition of heat will cause the reaction to shift in the direction which is endothermic, because that direction will use up heat (relieve the stress).

7. The characteristic that determines which direction the reaction shifts is whether the reaction is endothermic or exothermic.

8. The **production** of ammonia results in a decrease in the total number of moles in a closed system and therefore a **decrease** in the total pressure of the system. When the pressure of an ammonia system is increased, the equilibrium will shift to **decrease** the total pressure and will cause the **production** of ammonia to relieve the stress.

9. No. For a system containing gases, an increase in pressure will shift the equilibrium in the direction of the fewest number of gas molecules, thereby decreasing the total number of gas molecules and causing a decrease in pressure which relieves the stress (extra pressure). A change in pressure will not affect systems not containing gases because pressure does not affect solids, liquids, or aqueous substances.

10. If the reaction has unequal numbers of reactant and product gas molecules, the equilibrium can shift. The equilibrium will shift to the side with fewer gas molecules when there is an increase in pressure. Again, if there are no gas molecules involved in the equilibrium, changes in pressure will not affect the equilibrium.

11. Because the total number of molecules are the same on both sides of the reaction equation the stress (increased pressure) cannot be relieved by shifting one direction or the other.

12. When studying one characteristic in an investigation, it is good scientific procedure to keep all other factors constant so you can be sure that only the manipulated factor is causing the change.

13. When extra reactants are added to the ammonia system at equilibrium, the system will relieve the stress by **producing** ammonia. When extra ammonia is added to the system, the stress is relieved by **decomposing** ammonia. The equilibrium shifts to **use up** the added substance.

14. Controlling equilibria can make industrial processes more simple, more reliable, and more economical.

15.

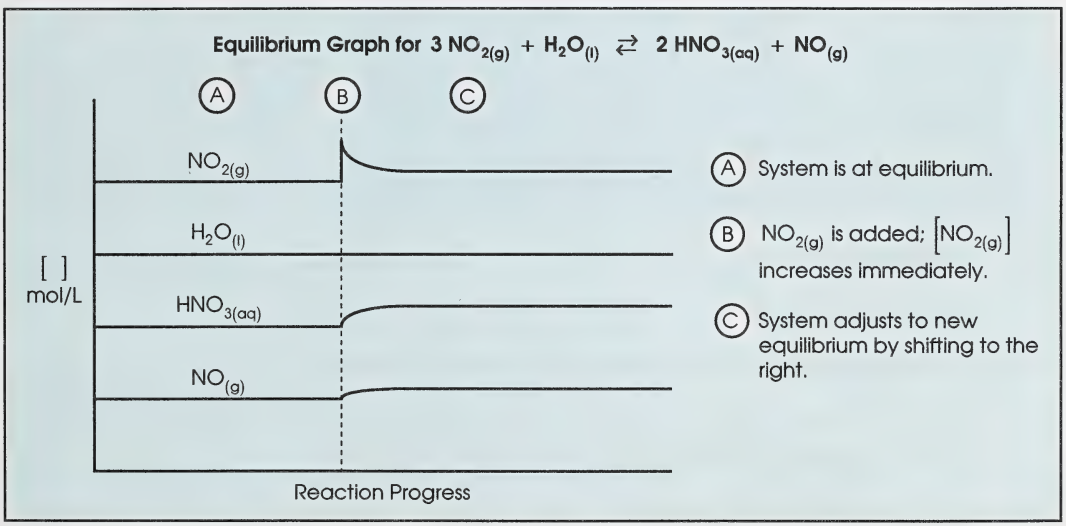
16.

Stress	Direction Equilibrium Shifts
Add reactants	**to right (products)**
Add products	**to left (reactants)**

17. a. Increasing the $\left[NO_{2(g)}\right]$ and decreasing the $\left[HNO_{3(aq)}\right]$ will cause a shift in the equilibrium towards the right.

Note: You might be inclined to say that increasing the $\left[H_2O_{(l)}\right]$ will shift the equilibrium towards the right. However, you cannot change the **concentration** of water – it is fixed by its density.

b.



18. The addition of heat favours the **products** in an endothermic reaction and the **reactants** in an exothermic reaction. Lowering the temperature of a reaction shifts an equilibrium to **products** if the reaction is exothermic and shifts an equilibrium to reactants in an **endothermic** process.

19. The reaction is endothermic because it shifts to the right (to produce products) to use up heat. Therefore, the heat term in the equation would be on the left side.

20. Since gas bubbles are produced, the system has shifted to the right. The system, when cooled, will shift the equilibrium to produce energy. Therefore, the system has shifted to the right to produce energy and must be exothermic.

21. a. decrease
b. increase
c. decrease or relieve
d. decreasing
e. increase
f. increasing
g. different
h. smaller or lesser
i. greater or higher

22. A change in pressure has no effect because the total number of gas molecules cannot be changed by shifting in either direction.

23. a. A catalyst has no net effect on a system already at equilibrium.

b. A catalyst will help the system get to equilibrium faster, but the position of the equilibrium will be the same whether or not the catalyst is present.

24.

Applied Stress	Response of System
Raise the temperature of an exothermic equilibrium system.	**shifts to reactants to use up energy**
Compress a gaseous system that has more reactant gases than product gases.	**shifts to products to reduce pressure**
Add a catalyst to a system at equilibrium.	**no shift occurs**
Add more of a reactant.	**shifts to products to use up reactant**
Lower the temperature of an endothermic equilibrium system.	**shifts to reactants to produce energy**
Remove products from an equilibrium system.	**shifts to products to add products**
Add more solute to a saturated solution.	**no change; maximum [ ] of ions already**
Increase the volume of a gaseous equilibrium system.	**shifts to side with most gas molecules**
Add a catalyst to a system not yet at equilibrium.	**gets to equilibrium faster**

25. a. An increase in $\left[CO_{2(g)}\right]$ will shift Equilibrium (1) to the right, causing the production of more precipitate.

b. A decrease in pressure should cause Equilibrium (2) to shift left to create more $CO_{2(g)}$ and increase the pressure. This will cause the HBb indicator to turn blue, which reflects the decrease in $\left[H^{+}_{(aq)}\right]$. An increase in pressure will cause the opposite effects.

26. A sample set of evidence is as follows.

- The entire solution turned blue when HBb was added to the limewater.
- When soda water was initially added, a milky precipitate formed in the blue solution.
- Just before the solution started to change colour, the precipitate disappeared.
- The solution changed colour from blue to yellow-green after the addition of about 23 drops of soda water.
- In the syringe, as pressure was reduced (the plunger drawn back), bubbles formed and the solution changed colour from yellow-green to blue.
- When the pressure was increased again (the plunger pushed forward), the solution changed colour from blue to yellow-green and any bubbles evident decreased in size.

27. Because the predictions made were based on Le Châtelier's principle, and the results confirmed the prediction, Le Châtelier's principle was confirmed.

- The formation of a precipitate with the addition of soda water to limewater confirmed the Equilibrium (1) shift to the right.
- The change in colour of the syringe solution from yellow-green to blue and vice versa confirmed the decrease, and subsequent increase, in the $\left[H^{+}_{(aq)}\right]$. The formation of bubbles with a decrease in pressure would also confirm the $CO_{2(g)}$ coming out of solution.

28. The predictions according to Le Châtelier's principle were consistently confirmed by the experimental results. This leads to the judgement that the design and procedure of the investigation was adequate. The procedure was easy to perform and leads one to be confident in the results. A white background and light source would help in making the observations.

When the limewater was added, the pressure and temperature were kept constant. When the pressure was changed, the temperature and amounts of substances in the system were kept constant.

29. Applying Le Châtelier's principle, an increase in the temperature of the cobalt(II) ion equilibrium should shift the equilibrium to the left causing a bluer colour; decreasing the temperature should shift the equilibrium to the right, causing a pinker colour. The addition of water should cause the equilibrium to shift to the right (more pink) to use up the water. An increase in the chloride ion concentration (from the sodium chloride) will shift the equilibrium to the left (bluer). The addition of $AgNO_{3(aq)}$ should cause the precipitation of $AgCl_{(s)}$, thus removing $Cl^{-}_{(aq)}$ from the system. This should cause an equilibrium shift to the right to try to replenish the $Cl^{-}_{(aq)}$ ions and cause a pinker colour. This shift to the right will partially offset the changes (stresses) made to the chemical system.

30. A suitable procedure might be as follows:

- Place about 2 mL of the cobalt(II) chloride equilibrium mixture in each of two small test tubes.
- Place one of the test tubes in a hot-water bath and the other in an ice-water bath; record any changes. Discard the solutions into a waste container.
- Place about 2 mL of the equilibrium mixture into a clean **dry** test tube and add a few drops of water. Record any changes. Discard the solution into a waste container.
- Place about 2 mL of the equilibrium mixture in each of two clean test tubes. Add a few grains of solid sodium chloride to one of the tubes. Add a few (10-15) drops of 0.10 mol/L silver nitrate solution to the other tube.
  - Stopper and shake each tube. Add more materials if necessary and shake again.
  - Record any observations. Discard the solutions into a waste container.

31. The following is a sample data table.

Applied Stress	Observations
hot water	turned bluer
cold water	turned pinker
added water	turned pinker
added $NaCl_{(s)}$	turned bluer
added $AgNO_{3(aq)}$	precipitate formed; solution turned pink

32.

Applied Stress	Predicted Change	Observed Change
addition of heat	**turn bluer**	**turned bluer**
reduction of heat	**turn pinker**	**turned pinker**
addition of water	**turn pinker**	**turned pinker**
addition of $NaCl_{(s)}$	**turn bluer**	**turned bluer**
addition of $AgNO_{3(aq)}$	**turn pinker**	**turned pinker**

33. Le Châtelier's principle did successfully predict the outcomes of the changes to the equilibrium system. As shown in the preceding question, all outcomes agreed with the predictions.

34. The addition of silver nitrate solution causes the chloride ion concentration to go down because the silver and chloride ions are insoluble together and they fall out of solution as the precipitate. This shifts the equilibrium to the right.

35. The experimental design was adequate as it provided an answer to the problem and was easy to perform. Using the same equilibrium mixture and the same volume was important to ensure consistency. The predictions were verified with a high degree of certainty.

36. **Textbook question 12:**

The three types of changes that shift the position of a chemical equilibrium are changes in concentration, temperature, and pressure.

Textbook question 13:

a. The equilibrium will shift to the right, producing more $H_2O_{(g)}$.

b. The equilibrium will shift to the left, reducing the $\left[H^+_{(aq)}\right]$ and $\left[OH^-_{(aq)}\right]$; however, the final $\left[OH^-_{(aq)}\right]$ is still higher than the original concentration because the system does not totally get rid of the amount of $OH^-_{(aq)}$ added.

c. The equilibrium shifts to the right, producing more $CaO_{(s)}$ and $CO_{2(g)}$. You CANNOT say that the *concentration* of $CaO_{(s)}$ is increased because the concentration of a solid does not change – it is fixed by its density.

d. The equilibrium shifts to the right, increasing the $\left[H^{+}_{(aq)}\right]$ and the $\left[CH_3COO^{-}_{(aq)}\right]$.

Textbook question 14:

a. The equilibrium shifts to the right, producing more $NO_{(g)}$ and $H_2O_{(g)}$.

b. The equilibrium shifts to the left, consuming $NO_{(g)}$, $H_2O_{(g)}$, and energy.

c. The equilibrium shifts to the right, producing more $NO_{(g)}$ and $H_2O_{(g)}$.

d. The equilibrium shifts to the left, consuming $NO_{(g)}$ and $H_2O_{(g)}$ to make fewer molecules.

Textbook question 15:

To increase the production of ammonia, you would need to shift the equilibrium to the right. Therefore, you could lower the temperature, raise the pressure, add $N_{2(g)}$ or $H_{2(g)}$, or remove $NH_{3(g)}$.

Textbook question 16:

a. Assuming that the left test tube is more yellow, the equilibrium favours reactants.

b. As more water is added to the system, the equilibrium shifts to the right, increasing the concentration of $Cu(H_2O)_4^{2+}{}_{(aq)}$ which is blue.

c. Obtain two test tubes of the $CuCl_4^{2-}{}_{(aq)} - Cu(H_2O)_4^{2+}{}_{(aq)}$ equilibrium mixture. Place one test tube into an ice-water bath and the other test tube into a hot-water bath. If the test tube in the ice-water bath burns blue, and the hot-water bath test tube turns yellowish green, then the reaction is exothermic. If the opposite happens, the reaction is endothermic. (Only one water bath is necessary; the other is more for confirmation.)

37. The First World War prompted the development of the Haber process. Germany needed the nitrogen compounds to make ammunition and the British navy was blockading its source (Chile).

38. Three stresses that could cause the process to favour the production of ammonia would be higher pressures (more reactant gases than product), lower temperatures (since the reaction is exothermic), and removing ammonia (reaction shifts to create more).

39. a. The reaction was too slow at lower temperatures; some favourability was given up to have the reaction go more quickly.

    b. The use of an iron oxide catalyst made it possible to do the reaction at lower temperatures.

40. Because Haber's discovery was being used to aid the Nazi war effort, the international community protested the award.

Section 1: Follow-up Activities

Extra Help

1. a. vi c. vii e. ix g. x i. iii
   b. iv d. v f. viii h. ii j. i

2.

Stress	Direction of Shift	Change in $[CO_{(g)}]$
Decrease the volume of the container.	**no shift**	**All components increase in [ ], including CO, because of the volume decrease.**
Add heat.	**right**	**There is an increase since the reaction is endothermic.**
Remove $H_2O_{(g)}$ using a desiccator.	**right**	**There is an increase; the reaction will shift to produce $CO_{(g)}$.**
Add $CO_{2(g)}$.	**right**	**There is an increase; the reaction will try to use up excess reactant.**
Put the reaction vessel in an ice bath.	**left**	**There is a decrease; the reaction will shift to produce heat (to the reactants).**
Add a catalyst.	**no shift**	**no change**

3. a. The equilibrium law for this reaction is $K = \dfrac{[CO_{(g)}][H_2O_{(g)}]}{[H_{2(g)}][CO_{2(g)}]}$

   b.

Reaction	$[CO_{2(g)}]$ (mol/L)	$[H_{2(g)}]$ (mol/L)	$[CO_{(g)}]$ (mol/L)	$[H_2O_{(g)}]$ (mol/L)
1	1.8	2.0	**0.60**	1.2
2	0.5	**6**	0.3	2.0
3	8.0	14	35	**0.64**

A sample calculation (for Reaction 1) is as follows:

$$\left[CO_{(g)}\right] = \frac{\left[H_{2(g)}\right]\left[CO_{2(g)}\right]K}{\left[H_2O_{(g)}\right]}$$

$$= \frac{(2.0 \text{ mol/L})(1.8 \text{ mol/L})0.20}{(1.2 \text{ mol/L})}$$

$= 0.60$ mol/L The concentration of $CO_{(g)}$ is 0.60 mol/L.

4. a. You know that equilibrium had been reached because the observable property (colour) stopped changing.

b. The percent reaction is calculated as follows:

amount of $HI_{(g)}$ reacted = original amount – final amount

$= 4.0 \text{ mol} - 1.2 \text{ mol}$

$= 2.8$ mol

maximum amount of $HI_{(g)}$ that could have reacted $= 4.0$ mol

$$\% \text{ reaction} = \frac{2.8 \text{ mol}}{4.0 \text{ mol}} \times 100\%$$

$= 70\%$

The percent reaction is 70%.

c. The reaction favours products since >50% of the $HI_{(g)}$ reacted.

d. The value of K is found as follows:

Step 1

$$K = \frac{\left[H_{2(g)}\right]\left[I_{2(g)}\right]}{\left[HI_{(g)}\right]^2}$$

Step 2

$$\text{initial } \left[HI_{(g)}\right] = \frac{4.0 \text{ mol}}{1.0 \text{ L}}$$

$= 4.0$ mol/L

$$\text{equilibrium } \left[HI_{(g)}\right] = \frac{1.2 \text{ mol}}{1.0 \text{ L}}$$

$= 1.2$ mol/L

Step 3

	$2\,HI_{(g)}$	$\rightleftarrows$	$H_{2(g)}$	+	$I_{2(g)}$
initial concentrations (mol/L)	4.0		0.0		0.0
changed concentrations (mol/L)	−2.8		$\frac{1}{2}(2.8)$		$\frac{1}{2}(2.8)$
equilibrium concentrations (mol/L)	1.2		1.4		1.4

Step 4

$$K = \frac{\left[H_{2(g)}\right]\left[I_{2(g)}\right]}{\left[HI_{(g)}\right]^2}$$

$$= \frac{(1.4\ \text{mol/L})(1.4\ \text{mol/L})}{(1.2\ \text{mol/L})^2}$$

$$= 1.4$$

The equilibrium constant is 1.4.

Enrichment

1. Investigation: K_{sp} of Calcium Hydroxide

Experimental Design

A sample of saturated calcium hydroxide will be titrated with a dilute solution of hydrochloric acid using bromothymol blue indicator.

The hydrochloric acid will react in a 2:1 ratio with the calcium hydroxide solution; the concentration of each ion at equilibrium can be calculated and the K_{sp} found.

Materials

- 20 mL of $Ca(OH)_{2(aq)}$ (saturated and filtered)
- 40 mL of 0.050 mol/L $HCl_{(aq)}$
- bromothymol blue indicator
- 10 mL pipet and bulb
- thermometer
- 250 mL Erlenmeyer flask
- titrating equipment (buret, stand, clamp)

Procedure

- Pipet 10 mL of $Ca(OH)_{2(aq)}$ into a clean Erlenmeyer flask and add a few drops of indicator. Record the colour and temperature of the solution.
- Rinse and fill a buret with standardized $HCl_{(aq)}$ (0.050 mol/L). Record the initial buret reading.
- Titrate the $Ca(OH)_{2(aq)}$ with the $HCl_{(aq)}$ to a visible endpoint. Record the final buret reading.
- Repeat the procedure until the results of the titrations are reasonably close to each other.

Evidence

	Trial 1	Trial 2
volume of $Ca(OH)_{2(aq)}$ (mL)	10.0	10.0
initial buret reading of $HCl_{(aq)}$ (mL)	2.50	6.60
final buret reading of $HCl_{(aq)}$ (mL)	6.60	10.68
temperature of $HCl_{(aq)}$ (°C)	21.0	21.0
endpoint colour (bromothymol blue)	green	green

Analysis

Trial 1:

volume of $HCl_{(aq)}$ used

$= 6.60\text{ mL} - 2.50\text{ mL}$

$= 4.10\text{ mL}$

Trial 2:

volume of $HCl_{(aq)}$ used

$= 10.68\text{ mL} - 6.60\text{ mL}$

$= 4.08\text{ mL}$

Average volume of $HCl_{(aq)}$ used

$$= \frac{4.10\text{ mL} + 4.08\text{ mL}}{2}$$

$= 4.09\text{ mL}$

$$Ca(OH)_{2(aq)} + 2\,HCl_{(aq)} \rightleftarrows CaCl_{2(aq)} + 2\,H_2O_{(l)}$$

10.0 mL	0.050 mol/L
C = ?	4.09 mL

$$C_{(Ca(OH)_2)} = 4.09\ \cancel{mL\ HCl} \times \frac{0.050\ \cancel{mol\ HCl}}{1\ \cancel{L\ HCl}} \times \frac{1\text{ mol }\cancel{Ca(OH)_2}}{2\ \cancel{mol\ HCl}} \times \frac{1}{10.0\ \cancel{mL}\ \cancel{Ca(OH)_2}}$$

$$= 1.023 \times 10^{-2}\text{ mol/L}$$

$$= 1.0 \times 10^{-2}\text{ mol/L}$$

$$Ca(OH)_{2(aq)} \rightleftarrows Ca^{2+}_{(aq)} + 2\,OH^-_{(aq)}$$

1.023×10^{-2} mol/L [ ] = ? [ ] = ?

$$\left[Ca^{2+}_{(aq)}\right] = \left[Ca(OH)_{2(aq)}\right]$$

$$= 1.023 \times 10^{-2}\text{ mol/L}$$

$$= 1.0 \times 10^{-2}\text{ mol/L}$$

$$\left[OH^-_{(aq)}\right] = 2 \times \left[Ca(OH)_{2(aq)}\right]$$

$$= 2\left(1.023 \times 10^{-2}\right)$$

$$= 2.045 \times 10^{-2}\text{ mol/L}$$

$$= 2.0 \times 10^{-2}\text{ mol/L}$$

$$Ca(OH)_{2(s)} \rightleftarrows Ca^{2+}_{(aq)} + 2\,OH^-_{(aq)}$$

$$K_{sp} = [Ca^{2+}_{(aq)}][OH^-_{(aq)}]^2$$
$$= (1.023\times10^{-2}\ mol/L)(2.045\times10^{-2}\ mol/L)^2$$
$$= 4.278\times10^{-6} \quad \text{(Remember, units are not necessary.)}$$
$$= 4.3\times10^{-6} \quad \text{The } K_{sp} \text{ of } Ca(OH)_{2(s)} \text{ is } 4.3\times10^{-6}.$$

Evaluation

Confidence could be shown in the results. The procedure seemed to work well. Errors could have been made in titrations. Care was taken to have no air in the tip, and the buret and pipet were rinsed with the appropriate solutions. The indicator acted as a cue to stop titrating. Perhaps using a pH meter might have produced better results.

2. When a low temperature favours product formation, performing the reaction at a low temperature sometimes makes the reaction too slow. The equilibrium constant only tells you what the final conditions will be, but nothing about the rate of the reaction. A balance must be achieved between favourable equilibrium conditions and speed of production.

3. **Textbook question 35:**

The reaction to produce wolfram(II) iodide, $W_{(s)} + I_{2(g)} \rightleftarrows WI_{2(g)}$, is exothermic; therefore, a high temperature of the system forces the equilibrium to shift to form reactants. Tungsten gets deposited on the filament. The iodine establishes an equilibrium with wolfram and wolfram(II) iodide, which at a high temperature favours the filament.

Textbook question 36:

As in Investigation: Equilibrium Shifts, reducing the air pressure reduces the quantity of oxygen that can be dissolved by the blood. At high altitudes, the body increases the number of red blood cells (containing haemoglobin) in the blood so that the blood can take up more oxygen. Because it takes some time (about a month) to make enough red blood cells to overcome the loss of oxygen due to lack of pressure, athletes train at a high altitude location prior to a big event to allow their bodies time to adjust.

4. ## Investigation 14.5: Studying a Chemical Equilibrium System

Problem

What is the effect of an energy change on the following equilibrium system?

$$Fe^{3+}_{(aq)} + SCN^-_{(aq)} \rightleftarrows FeSCN^{2+}_{(aq)}$$

Prediction

Forming bonds is an exothermic process; therefore, this reaction is most likely exothermic. In the cool-water bath, a darker colour should be noted; in a hot bath a lighter colour should be seen.

Experimental Design

To determine if the iron(III) thiocyanate equilibrium is endothermic or exothermic, the equilibrium mixture is separated into two samples. One sample is placed in a cold-water bath; the other sample is placed in a hot-water bath. Colour changes are observed in each case. A darker colour indicates more $FeSCN^{2+}_{(aq)}$ (a shift to the right); a lighter colour indicates less $FeSCN^{2+}_{(aq)}$ (a shift to the left).

Materials

- 0.2 mol/L $Fe(NO_3)_{3(aq)}$
- 0.2 mol/L $KSCN_{(aq)}$
- 2 test tubes with rack
- 100 mL beaker
- crushed ice and hot water
- safety glasses and lab apron

Procedure

- The equilibrium mixture is formed by mixing 5 mL of 0.2 mol/L iron(III) nitrate solution with 5 mL of 0.2 mol/L potassium thiocyanate solution and diluting to a 1 L volume.
- Place two equal-sized samples of the equilibrium mixture (about 5 mL) in separate test tubes.
- Place each test tube in a different water bath – one hot, one cold.
- Observe and record any changes that occur.

Evidence

- The test-tube solution in the hot-water bath became lighter in colour.
- The test-tube solution in the cold-water bath became darker in colour.

Analysis

Since a decrease in temperature would cause the equilibrium to shift to produce more heat, according to Le Châtelier's principle, and the colour got darker in the cold test tube, indicating more $FeSCN^{2+}_{(aq)}$ being produced, the forward reaction is exothermic. This is confirmed by the decrease in colour in the hot-water test tube, which will work in the endothermic direction to use up heat. The lightening of colour is due to the $FeSCN^{2+}_{(aq)}$ being used up.

Evaluation

The experimental design seemed adequate since it provided an answer to the problem. The prediction is judged to be verified since it agreed with the results. The design and procedure seemed adequate since they provided consistent results and were easy to carry out.

Section 2: Activity 1

1. Measuring the degree of conductivity is not a suitable diagnostic test because acidic, basic, and neutral substances could all conduct electricity in aqueous solution. (**Note:** Neutral means not acidic or basic, but could be ionic and thus an electrolyte.)

2. A suitable diagnostic test would be to test each of the acidic, basic, and neutral solutions with both blue and red litmus paper. The following results would be obtained:

Solution	Colour of Blue Litmus	Colour of Red Litmus
acidic	red	red
basic	blue	blue
neutral	blue	red

3. Dissociation refers to ionic compounds separating into ions in solution whereas ionization is the reaction of neutral atoms or molecules to form charged species (ions).

4.

Definition	Acids	Bases
Theoretical (Arrhenius)	• dissolve in water and ionize to increase the $[H^+_{(aq)}]$	• dissolve in water and dissociate to increase the $[OH^-_{(aq)}]$
Empirical	• are electrolytes • turn blue litmus red • neutralize bases • react with active metals to produce $H_{2(g)}$	• are electrolytes • turn red litmus blue • neutralize acids • feel slippery to the touch

5. Tasting is not suitable in the laboratory because many substances may be harmful when eaten. Tasting is included in the definition because many foods that are safe to eat have been shown by other tests to be acidic or basic.

6. a. Since conductivity is due to the presence of ions, there must be ions present in pure water.

   b. Since the conductivity is so low, there are very few ions present.

   c. $H_2O_{(l)} \overset{<10^{-6}\%}{\rightleftarrows} H^+_{(aq)} + OH^-_{(aq)}$

   d. The reactants (water molecules) are greatly favoured in this water equilibrium.

7. a. $K_w = [H^+_{(aq)}][OH^-_{(aq)}]$

   b. Because water is the solvent and predominant substance in most aqueous solutions, the $[H_2O_{(l)}]$ stays relatively constant as compared to the other quantities (ions, solute molecules, etc.). This constant concentration is included as part of the equilibrium constant, K_w, value for this system.

8. a. Since K_w is very small, the reaction greatly favours reactants and there are not very many ions formed.

   b. Yes, because the low conductivity of pure water indicates few ions are present.

9. **Textbook question 17:**

$$K_w = [H^+(aq)][OH^-(aq)]$$

$$[OH^-(aq)] = \frac{1.0 \times 10^{-14}\ (mol/L)^{\cancel{2}}}{4.40 \times 10^{-3}\ \cancel{mol/L}}$$

$$= 2.3 \times 10^{-12}\ mol/L$$

The $[OH^-(aq)]$ is 2.3×10^{-12} mol/L.

Textbook question 18:

$$K_w = [H^+(aq)][OH^-(aq)]$$

$$[H^+(aq)] = \frac{1.0 \times 10^{-14}\ (mol/L)^{\cancel{2}}}{2.99 \times 10^{-4}\ \cancel{mol/L}}$$

$$= 3.3 \times 10^{-11}\ mol/L$$

The $[H^+(aq)]$ is 3.3×10^{-11} mol/L.

Textbook question 19:

$$n_{(HCl)} = 0.37\ \cancel{g} \times \frac{1\ mol}{36.46\ \cancel{g}}$$

$$= 0.010\ 15\ mol$$

$$HCl_{(aq)} \rightleftarrows H^+(aq) + Cl^-(aq)$$

$$n_{(H^+)} = 0.010\ 15\ \cancel{mol\ HCl} \times \frac{1\ mol\ H^+}{1\ \cancel{mol\ HCl}}$$

$$= 0.010\ 15\ mol\ H^+(aq)$$

$$[H^+(aq)] = \frac{0.010\ 15\ mol}{0.250\ L}$$

$$= 0.040\ 59\ mol/L$$

$$K_w = [H^+(aq)][OH^-(aq)]$$

$$[OH^-(aq)] = \frac{1.0 \times 10^{-14}\ (mol/L)^{\cancel{2}}}{0.040\ 59\ \cancel{mol/L}}$$

$$= 2.5 \times 10^{-13}\ mol/L$$

The $[OH^-(aq)]$ is 2.5×10^{-13} mol/L.

Textbook question 20: (page 445)

$$Ca(OH)_{2(s)} \rightleftarrows Ca^{2+}(aq) + 2\ OH^-(aq)$$

$$[OH^-(aq)] = \frac{6.9\ m\cancel{mol\ Ca(OH)_2}}{L} \times \frac{2\ mol\ OH^-}{1\ \cancel{mol\ Ca(OH)_2}}$$

$$= \frac{13.8\ mmol\ OH^-}{L}$$

$$= 1.38 \times 10^{-2}\ mol/L\ OH^-$$

$$K_w = [H^+(aq)][OH^-(aq)]$$

$$[H^+(aq)] = \frac{1.0 \times 10^{-14}\ (mol/L)^{\cancel{2}}}{1.38 \times 10^{-2}\ \cancel{mol/L}}$$

$$= 7.2 \times 10^{-13}\ mol/L$$

The $[H^+(aq)]$ is 7.2×10^{-13} mol/L.

Textbook question 21:

$$KOH_{(s)} \rightleftarrows K^+_{(aq)} + OH^-_{(aq)}$$

$$n_{(KOH)} = 20.0\ \cancel{g} \times \frac{1\ mol}{56.11\ \cancel{g}}$$
$$= 0.356\ 44\ mol$$

$$n_{(OH^-)} = 0.356\ 44\ \cancel{mol\ KOH} \times \frac{1\ mol\ OH^-}{1\ \cancel{mol\ KOH}}$$
$$= 0.356\ 44\ mol\ OH^-$$

$$[OH^-_{(aq)}] = \frac{0.356\ 44\ mol}{0.500\ L}$$
$$= 0.712\ 89\ mol/L$$

$$K_w = [H^+_{(aq)}][OH^-_{(aq)}]$$
$$[H^+_{(aq)}] = \frac{1.0 \times 10^{-14}\ (mol/L)^{\cancel{2}}}{0.712\ 89\ \cancel{mol/L}}$$
$$= 1.4 \times 10^{-14}\ mol/L$$

The $[H^+_{(aq)}]$ is 1.4×10^{-14} mol/L.

Textbook question 22:

$$H_2O_{(l)} \rightleftarrows H^+_{(aq)} + OH^-_{(aq)}$$

$$[H_2O_{(l)}] = 1000\ \cancel{g} \times \frac{1\ mol}{18.02\ \cancel{g}} \times \frac{1}{1.00\ L}$$
$$= 55.494\ mol/L$$

$$\%\ ionization = \frac{[H^+_{(aq)}]}{[H_2O_{(l)}]} \times 100\%$$
$$= \frac{1.0 \times 10^{-7}\ \cancel{mol/L}}{55.494\ \cancel{mol/L}} \times 100\%$$
$$= 1.8 \times 10^{-7}\%$$

The percent ionization of water is 1.8×10^{-7}%.

Textbook question 20: (page 457)

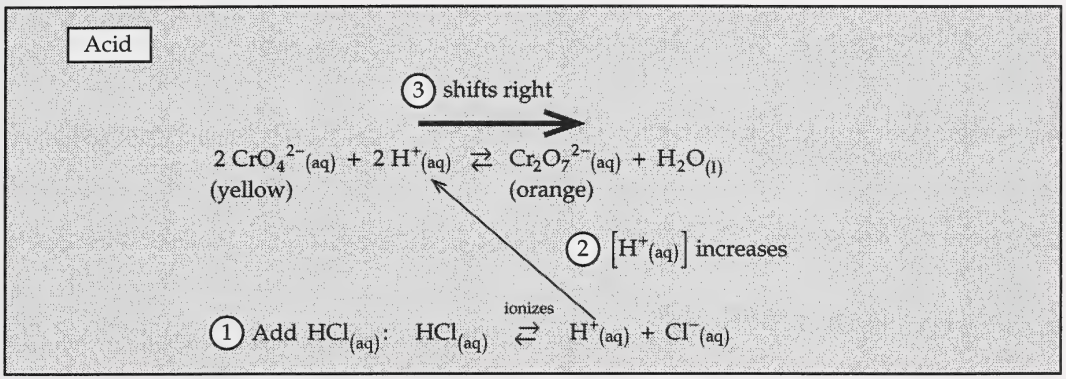
$$[H^+_{(aq)}] = 1.3\ mmol/L$$
$$= 1.3 \times 10^{-3}\ mol/L$$

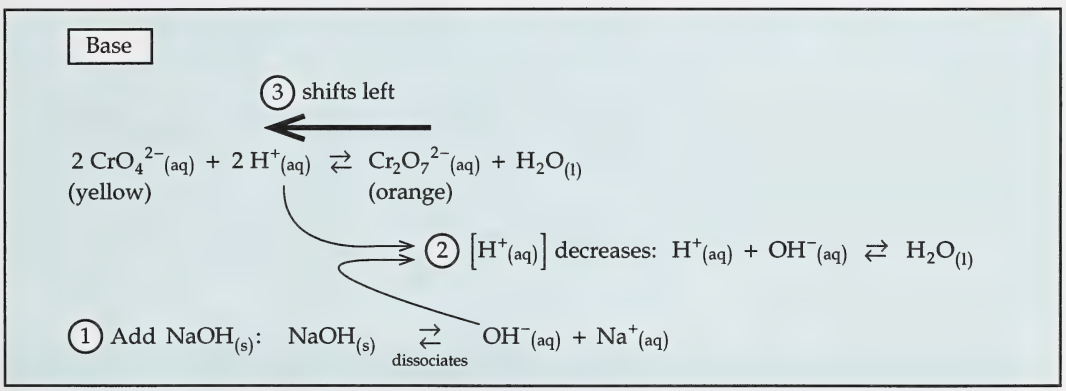
$$K_w = [H^+_{(aq)}][OH^-_{(aq)}]$$
$$[OH^-_{(aq)}] = \frac{1.0 \times 10^{-14}\ (mol/L)^{\cancel{2}}}{1.3 \times 10^{-3}\ \cancel{mol/L}}$$
$$= 7.7 \times 10^{-12}\ mol/L$$

The $[OH^-_{(aq)}]$ is 7.7×10^{-12} mol/L.

10. a. Your prediction may be as follows:

Based on Le Châtelier's principle, when the $[H^+_{(aq)}]$ is increased in the chromate-dichromate equilibrium, the equilibrium will shift to the right toward product formation, increasing the $[Cr_2O_7^{2-}{}_{(aq)}]$ and decreasing the $[CrO_4^{2-}{}_{(aq)}]$. Increasing the $[OH^-_{(aq)}]$ will achieve the opposite result.




Base
3 shifts left
2 CrO4 2−(aq) + 2 H+(aq) ⇄ Cr2O7 2−(aq) + H2O(l)
(yellow)
(orange)
2 [H+(aq)] decreases: H+(aq) + OH−(aq) ⇄ H2O(l)
1 Add NaOH(s): NaOH(s) ⇄ OH−(aq) + Na+(aq)
dissociates

b. Your experimental design may be as follows:

Equal volumes of the chromate-dichromate will be separately tested with $HCl_{(aq)}$ and $NaOH_{(s)}$. Their colours will be compared to another sample of the original chromate-dichromate equilibrium. An increase in orange colour due to increased dichromate ions indicates a shift to the right (products). A decrease in the orange colour (an increase in yellow) will indicate a shift to the left (reactants).

Section 2: Activity 2

1. The number of significant digits communicated in the hydrogen ion concentration is the same as **the number of places after the decimal point in the pH value**.

2. If $pH = -\log[H^+]$, then $[H^+_{(aq)}] = 10^{-pH}$.

3.

pH	$[H^+_{(aq)}]$ (mol/L)	Classification (acidic/basic/neutral)
3.67	**2.1×10^{-4}**	**acidic**
4.611*	2.45×10^{-5}	**acidic**
10.1	**8×10^{-11}**	**basic**
11.0294*	9.345×10^{-12}	**basic**
5.63	**2.3×10^{-6}**	**acidic**

***Note:** In most cases, pH values are restricted to two decimal places.

Sample calculations:

Given a pH of 3.67:

$[H^+_{(aq)}] = \text{INV}\log(-\text{pH})$ **or** 10^{-pH}

$[H^+_{(aq)}] = \text{INV}\log(-3.67)$ **or** $10^{-3.67}$

$= 2.1\times10^{-4}$ mol/L

Given $[H^+_{(aq)}] = 2.45\times10^{-5}$ mol/L:

$pH = -\log[H^+_{(aq)}]$

$= -\log(2.45\times10^{-5})$

$= -(-4.611)$

$= 4.611$

4. Two very useful relationships to use when converting between pH, pOH, $[H^+_{(aq)}]$ and $[OH^-_{(aq)}]$ are

- pH + pOH = **14** at SATP
- $[H^+_{(aq)}] \times [OH^-_{(aq)}] = 1.0\times10^{-14}$ $(\text{mol/L})^2$ at SATP

5. **Textbook question 23:**

ACIDITY OF FOODS				
Food	$[H^+_{(aq)}]$ (mol/L)	$[OH^-_{(aq)}]$ (mol/L)	pH	pOH
oranges	5.5×10^{-3}	**1.8×10^{-12}**	**2.26**	**11.74**
asparagus	**4×10^{-9}**	**3×10^{-6}**	**8.4**	5.6
olives	**5.0×10^{-4}**	2.0×10^{-11}	**3.30**	**10.70**
blackberries	**4.0×10^{-4}**	**2.5×10^{-11}**	**3.40**	10.60

Sample calculations for asparagus: $pOH = 5.6$

$pH = 14 - 5.6$

$= 8.4$

$[H^+_{(aq)}] = 10^{-pH}$

$= 10^{-8.4}$

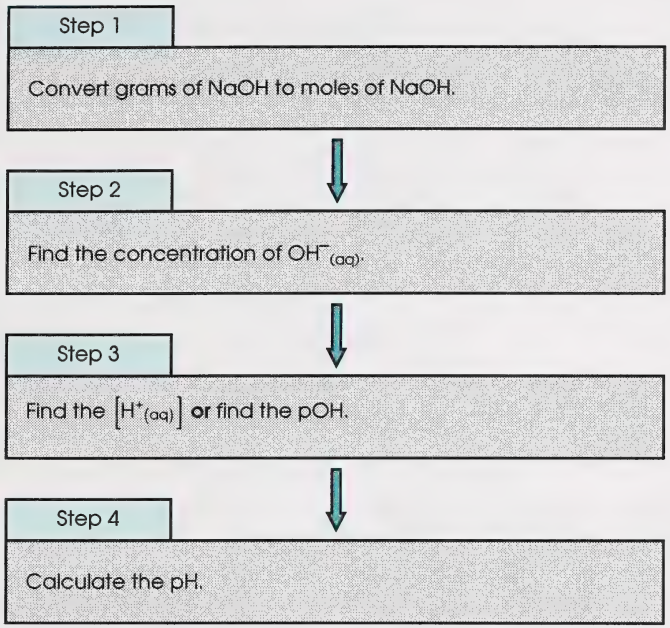
$= 4\times10^{-9}$ mol/L

$[OH^-_{(aq)}] = 10^{-pOH}$

$= 10^{-5.6}$

$= 3\times10^{-6}$ mol/L

6. a.



b. Variables

m = 2.4 g NaOH pH = ?
v = 250 mL solution

$$NaOH_{(s)} \rightleftarrows Na^+_{(aq)} + OH^-_{(aq)}$$

2.4 g [ ] = ?
250 mL

Step 1 Step 2

$$[OH^-_{(aq)}] = 2.4\ \cancel{g\ NaOH} \times \frac{1\ \cancel{mol\ NaOH}}{40.00\ \cancel{g\ NaOH}} \times \frac{1\ mol\ \cancel{OH^-}}{1\ \cancel{mol\ NaOH}} \times \frac{1}{0.250\ L\ \cancel{OH^-}}$$

$$= 0.24\ mol/L$$

Step 3

$$[H^+_{(aq)}] = \frac{K_w}{[OH^-_{(aq)}]}$$

$$= \frac{1.0 \times 10^{-14}\ (mol/L)^{\cancel{2}}}{0.24\ \cancel{mol/L}}$$

$$= 4.167 \times 10^{-14}\ mol/L$$

$$= 4.2 \times 10^{-14}\ mol/L$$

or

$$pOH = -\log[OH^-_{(aq)}]$$

$$= -\log(0.24)$$

$$= -(-0.6198)$$

$$= 0.62$$

Step 4

$$pH = -\log\left[H^+_{(aq)}\right]$$
$$= -\log\left(4.167 \times 10^{-14}\right)$$
$$= -(-13.3802)$$
$$= 13.38$$

or

$$pH = 14 - pOH$$
$$= 14 - 0.6198$$
$$= 13.3802$$
$$= 13.38$$

The pH of the solution is 13.38.

7. **Textbook question 24:**

Variables: $m = 26$ g NaOH pH = ?
$v = 150$ mL pOH = ?

$$NaOH_{(s)} \rightleftarrows Na^+_{(aq)} + OH^-_{(aq)}$$

$$\left[OH^-_{(aq)}\right] = 26\ \cancel{g\ NaOH} \times \frac{1\ \cancel{mol\ NaOH}}{40.00\ \cancel{g\ NaOH}} \times \frac{1\ mol\ \cancel{OH^-}}{1\ \cancel{mol\ NaOH}} \times \frac{1}{0.150\ L\ \cancel{OH^-}}$$
$$= 4.333\ mol/L$$
$$= 4.3\ mol/L$$

$$pOH = -\log\left[OH^-_{(aq)}\right]$$
$$= -\log(4.333)$$
$$= -0.6368$$
$$= -0.64$$

The pOH is -0.64.

$$pH = 14 - pOH$$
$$= 14 - (-0.6368)$$
$$= 14.6368$$
$$= 14.64$$

The pH is 14.64.

Textbook question 25:

Variables: $v = 500$ mL $m_{(KOH)} = ?$
pH = 11.5

$$KOH_{(s)} \rightleftarrows K^+_{(aq)} + OH^-_{(aq)}$$

Step 1

Find the $\left[H^+_{(aq)}\right]$ or pOH from the pH.

$$\left[H^+_{(aq)}\right] = 10^{-pH}$$
$$= 10^{-11.5} \text{ (or INV log } -11.5)$$
$$= 3.16 \times 10^{-12}\ mol/L$$
$$= 3 \times 10^{-12}\ mol/L$$

or

$$pOH = 14 - pH$$
$$= 14 - 11.5$$
$$= 2.5$$

Step 2 Since the required solute $\left(KOH_{(s)}\right)$ is an OH^--containing base, you need to find the $\left[OH^-_{(aq)}\right]$.

$$\left[OH^-_{(aq)}\right] = \frac{K_w}{\left[H^+_{(aq)}\right]}$$

$$= \frac{1.0 \times 10^{-14}\ (\text{mol/L})^2}{3.16 \times 10^{-12}\ \cancel{\text{mol/L}}}$$

$$= 3.16 \times 10^{-3}\ \text{mol/L}$$

$$= 3 \times 10^{-3}\ \text{mol/L}$$

or

$$\left[OH^-_{(aq)}\right] = 10^{-\text{pOH}}$$

$$= 10^{-2.5}$$

$$= 3.16 \times 10^{-3}\ \text{mol/L}$$

$$= 3 \times 10^{-3}\ \text{mol/L}$$

Step 3 Calculate the mass of $KOH_{(s)}$.

$$m_{(KOH)} = \frac{3.16 \times 10^{-3}\ \cancel{\text{mol OH}^-}}{\cancel{\text{L}}} \times 0.500\ \cancel{\text{L}} \times \frac{1\ \cancel{\text{mol KOH}}}{1\ \cancel{\text{mol OH}^-}} \times \frac{56.11\ \text{g}}{1\ \cancel{\text{mol KOH}}}$$

$$= 0.0887\ \text{g}$$

$$= 0.09\ \text{g}$$

The mass of $KOH_{(s)}$ is 0.09 g.

Textbook question 29:

Variables: $v = 2.00\ \text{L}$ $\quad m_{(NaOH)} = ?$

$\text{pH} = 10.35$

$$NaOH_{(aq)} \rightleftarrows Na^+_{(aq)} + OH^-_{(aq)}$$

$$\left[H^+_{(aq)}\right] = 10^{-\text{pH}}$$

$$= 10^{-10.35}$$

$$= 4.467 \times 10^{-11}\ \text{mol/L}$$

$$= 4.5 \times 10^{-11}\ \text{mol/L}$$

or

$$\text{pOH} = 14 - \text{pH}$$

$$= 14 - 10.35$$

$$= 3.65$$

$$\left[OH^-_{(aq)}\right] = \frac{K_w}{\left[H^+_{(aq)}\right]}$$

$$= \frac{1.0 \times 10^{-14}\ (\text{mol/L})^2}{4.467 \times 10^{-11}\ \cancel{\text{mol/L}}}$$

$$= 2.239 \times 10^{-4}\ \text{mol/L}$$

$$= 2.2 \times 10^{-4}\ \text{mol/L}$$

or

$$\left[OH^-_{(aq)}\right] = 10^{-\text{pOH}}$$

$$= 10^{-3.65}$$

$$= 2.239 \times 10^{-4}\ \text{mol/L}$$

$$= 2.2 \times 10^{-4}\ \text{mol/L}$$

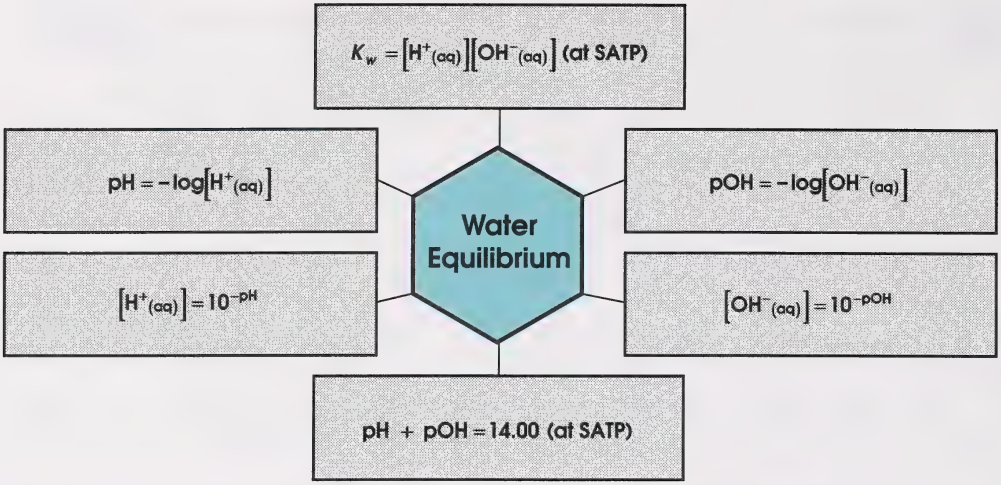
$$m_{(NaOH)} = \frac{2.239 \times 10^{-4}\ \cancel{\text{mol OH}^-}}{\cancel{\text{L}}} \times 2.00\ \cancel{\text{L}} \times \frac{1\ \cancel{\text{mol NaOH}}}{1\ \cancel{\text{mol OH}^-}} \times \frac{40.00\ \text{g}}{1\ \cancel{\text{mol NaOH}}}$$

$$= 0.01791\ \text{g}$$

$$= 0.018\ \text{g}$$

The mass of $NaOH_{(s)}$ dissolved is 0.018 g.

8.



9.

pH	$[H^+_{(aq)}]$ (mol/L)	Volume of Solution	Mass of HCl Needed
3.50	**3.2×10^{-4}**	250 mL	**2.9 mg**
2.611	2.45×10^{-3}	2.00 L	**0.179 g**
0.602	**2.50×10^{-1}**	400 mL	3.65 g
2.785	**1.64×10^{-3}**	3.50 L	**0.209 g**
3.00	**1.0×10^{-3}**	**2.0×10^{2} mL**	7.3 mg

Sample calculations:

$HCl_{(aq)} \rightleftarrows H^+_{(aq)} + Cl^-_{(aq)}$

Variables: pH = 3.50 v = 250 mL $[H^+_{(aq)}] = ?$ $m_{(HCl)} = ?$

$$[H^+_{(aq)}] = 10^{-pH}$$
$$= 10^{-3.50}$$
$$= 3.162 \times 10^{-4} \text{ mol/L}$$
$$= 3.2 \times 10^{-4} \text{ mol/L}$$

$$m_{(HCl)} = \frac{3.162 \times 10^{-4} \cancel{\text{mol H}^+}}{\cancel{\text{L}}} \times \frac{1 \text{ mol HCl}}{1 \cancel{\text{mol H}^+}} \times 0.250 \cancel{\text{L}} \times \frac{36.46 \text{ g}}{1 \cancel{\text{mol HCl}}}$$
$$= 0.002882 \text{ g}$$
$$= 2.9 \text{ mg}$$

10.

pH	pOH	$[OH^-_{(aq)}]$ (mol/L)	Volume	Mass of NaOH
10.35	**3.65**	**2.2×10^{-4}**	250 mL	**2.2 mg**
11.60	2.40	**4.0×10^{-3}**	**0.15 kL**	24 g
9.54	**4.46**	3.5×10^{-5}	**1.4 kL**	2.00 g
13.8	**0.2**	6×10^{-1}	3 L	**7×10^{1} g**
12.545	1.455	**3.51×10^{-2}**	**25.7 mL**	36.00 mg

Sample calculations:

$NaOH_{(s)} \rightleftarrows Na^+_{(aq)} + OH^-_{(aq)}$

Variables: $pOH = 2.40$ $m_{(NaOH)} = 24$ g $pH = ?$ $[OH^-_{(aq)}] = ?$ $v = ?$

$pH = 14 - pOH$
$= 14 - 2.40$
$= 11.60$

$[OH^-_{(aq)}] = 10^{-pOH}$
$= 10^{-2.40}$
$= 3.981\times10^{-3}$ mol/L
$= 4.0\times10^{-3}$ mol/L

$$v_{(OH^-)} = 24\ \cancel{\text{g NaOH}} \times \frac{1\ \cancel{\text{mol NaOH}}}{40.00\ \cancel{\text{g NaOH}}} \times \frac{1\ \cancel{\text{mol OH}^-}}{1\ \cancel{\text{mol NaOH}}} \times \frac{1\text{ L}}{3.981\times10^{-3}\ \cancel{\text{mol OH}^-}}$$
$$= 150.7\text{ L}$$
$$= 0.15\text{ kL}$$

11. A possible set of evidence is given in the following table.

Substance	pH	Substance	pH
vinegar	2.6	apple juice	3.8
soda water	3.9	toilet bowl cleanser	12.8
aspirin	3.5	dish soap	8.4
all-purpose cleaner	12.0	lemon juice	3.3

12. A possible set of evidence is in the following table.

Substance	Colour of pH paper	pH	Substance	Colour of pH paper	pH
vinegar	burnt orange	3	apple juice	yellow orange	4
soda water	yellow orange	4	toilet bowl cleanser	navy	13
aspirin	burnt orange	3	dish soap	green	8
all-purpose cleaner	indigo	12	lemon juice	burnt orange	3

13. The evidence collected seems to indicate that foods have a lower pH than substances used for cleaning. All foods tested had pH values <7; all cleaners (with the possible exception of vinegar) had pH values >7. Within the limitations of the substances tested, the trend seems to be that foods tend to be acidic and cleaners tend to be basic.

Section 2: Activity 3

1. Your experimental design may be as follows;

   Solutions of common acids with equal concentrations are tested for their acidity by measuring the pH of the solutions using a pH meter. The pH will act as a measure to put the acids in order from strong to weak. The manipulated variable is the acid type; the responding variable is the pH; the controlled variables are temperature, concentration, and equipment.

2. Your analysis may be as follows:

   By definition, a lower pH means a higher $\left[H^{+}_{(aq)}\right]$. From the evidence presented, the acids in order from lowest pH to highest pH (and therefore from highest acidity to lowest acidity) are $HCl_{(aq)}$ and $HNO_{3(aq)}$ (equal), $HF_{(aq)}$, $HCOOH_{(aq)}$, $CH_3COOH_{(aq)}$, and $HCN_{(aq)}$.

3. Other properties that could have been tested are as follows:

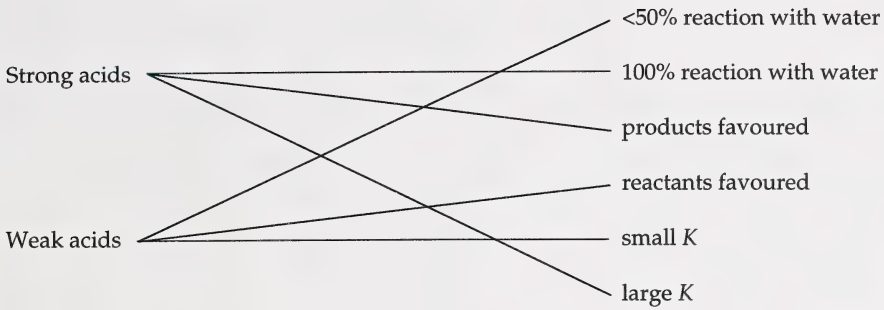
   - **conductivity –** The conductivity (as tested with a sensitive meter) would decrease in the same order as acid strength.
   - **rate of reaction with carbonates –** The rate of bubbling (how vigorous) would decrease in the same order as acid strength.
   - **rate of reaction with active metals –** The rate of reaction would decrease in the same order as acid strength.

4. Temperature and concentration changes might also cause the pH to change. Then you wouldn't know whether it was the concentration or temperature or strength that was creating the different pH values.

5.

Empirical Properties	Strong Acid (0.10 mol/L)	Weak Acid (0.10 mol/L)
conductivity	good conductor	poor conductor
pH	pH = 1 or lower	pH higher than 1
rate of reaction	fast reaction rates	slow reaction rates

6.



7. The equation when rearranged is as follows: $\left[H^{+}_{(aq)}\right] = \frac{\text{percent ionization}}{100} \times [\text{acid}]$

8. The calculations for the pH of $HCl_{(aq)}$ are as follows:

$$\begin{aligned} pH &= -\log\left[H^{+}_{(aq)}\right] \\ &= -\log(0.10) \\ &= 1.00 \end{aligned}$$

The calculations for the pH of $CH_3COOH_{(aq)}$ are as follows:

$$\begin{aligned} pH &= -\log\left[H^{+}_{(aq)}\right] \\ &= -\log(0.0013) \\ &= 2.89 \end{aligned}$$

9. a. There are six strong acids. They are as follows:

- perchloric acid, $HClO_{4(aq)}$
- hydroiodic acid, $HI_{(aq)}$
- hydrobromic acid, $HBr_{(aq)}$
- hydrochloric acid, $HCl_{(aq)}$
- sulphuric acid, $H_2SO_{4(aq)}$
- nitric acid, $HNO_{3(aq)}$

b.

Name of Acid	Chemical Formula	Percent Ionization	$\left[H^{+}_{(aq)}\right]$ (mol/L)	pH
hydrocyanic acid	$HCN_{(aq)}$	0.0078	7.8×10^{-6}	5.11
hydrofluoric acid	$HF_{(aq)}$	7.8	7.8×10^{-3}	2.11
phosphoric acid	$H_3PO_{4(aq)}$	23	2.3×10^{-2}	1.64
hypochlorous acid	$HClO_{(aq)}$	0.054	5.4×10^{-5}	4.27
hydrobromic acid	$HBr_{(aq)}$	100	0.10	1.00
nitrous acid	$HNO_{2(aq)}$	8.1	8.1×10^{-3}	2.09

Sample calculation for hydrocyanic acid – $HCN_{(aq)}$: % ionization from table – 0.0078%

$$\% \text{ ionization} = \frac{[H^+_{(aq)}]}{[\text{acid}]} \times 100$$

$$[H^+_{(aq)}] = \frac{\% \text{ ionization}}{100} \times [\text{acid}]$$

$$= \frac{0.0078}{100} \times 0.10 \text{ mol/L}$$

$$= 7.8 \times 10^{-6} \text{ mol/L}$$

$$\text{pH} = -\log[H^+_{(aq)}]$$

$$= -\log(7.8 \times 10^{-6})$$

$$= 5.11$$

10. You should have a flowchart that looks something like the following:

$$\text{pH} \xrightarrow[10^{-\text{pH}}]{} [H^+_{(aq)}] \xrightarrow[\frac{[H^+_{(aq)}]}{[\text{acid}]} \times 100]{} \% \text{ ionization}$$

11.

Acid (name and formula)	Acid Concentration (mol/L)	Measured pH	Percent Ionization
ethanoic, $CH_3COOH_{(aq)}$	0.15	2.78	**1.1**
ethanoic, $CH_3COOH_{(aq)}$	0.050	3.02	**1.9**
ethanoic, $CH_3COOH_{(aq)}$	0.010	3.37	**4.3**
hydrofluoric, $HF_{(aq)}$	0.15	2.00	**6.7**
hydrofluoric, $HF_{(aq)}$	0.050	2.26	**11**
hydrofluoric, $HF_{(aq)}$	0.010	2.64	**23**

A sample calculation for ethanoic acid $\left(CH_3COOH_{(aq)}\right)$ is as follows:

Variables: $\left[CH_3COOH_{(aq)}\right] = 0.15$ mol/L measured pH = 2.78 % ionization = ?

$$[H^+_{(aq)}] = 10^{-\text{pH}}$$

$$= 10^{-2.78}$$

$$= 1.660 \times 10^{-3} \text{ mol/L}$$

$$= 1.7 \times 10^{-3} \text{ mol/L}$$

$$\% \text{ ionization} = \frac{[H^+_{(aq)}]}{[\text{acid}]} \times 100$$

$$= \frac{1.660 \times 10^{-3} \text{ mol/L}}{0.15 \text{ mol/L}} \times 100$$

$$= 1.1\%$$

The percent ionization is 1.1%.

12. As the concentration of the acid decreased, the percent ionization increased.

Section 2: Activity 4

1. $HCN_{(aq)} \rightleftarrows H^+(aq) + CN^-(aq)$ $\quad K_a = \dfrac{[H^+(aq)][CN^-(aq)]}{[HCN_{(aq)}]}$

$HCOOH_{(aq)} \rightleftarrows H^+(aq) + HCOO^-(aq)$ $\quad K_a = \dfrac{[H^+(aq)][HCOO^-(aq)]}{[HCOOH_{(aq)}]}$

$H_3PO_{4(aq)} \rightleftarrows H^+(aq) + H_2PO_4^-(aq)$ $\quad K_a = \dfrac{[H^+(aq)][H_2PO_4^-(aq)]}{[H_3PO_{4(aq)}]}$

2. The K_a value is valid over a large range of concentrations, whereas the percent reaction or ionization is valid only at concentrations very close to the concentration at which it was originally calculated – 0.10 mol/L for your acid/base table values.

3. When the K_a value is at least 1000 (one thousand) times smaller than the concentration of the acid, within acceptable error (<5%), the concentration of the acid at equilibrium is the same as the original concentration.

4. $CH_3COOH_{(aq)} \rightleftarrows H^+(aq) + CH_3COO^-(aq)$

$$K_a = \frac{[H^+(aq)][CH_3COO^-(aq)]}{[CH_3COOH_{(aq)}]} = \frac{[H^+(aq)]^2}{[CH_3COOH_{(aq)}]}$$

$$[H^+(aq)] = \sqrt{K_a[\text{acid}]} = \sqrt{(1.8\times10^{-5}\ \text{mol/L})(0.15\ \text{mol/L})} = 1.643\times10^{-3}\ \text{mol/L} = 1.6\times10^{-3}\ \text{mol/L}$$

$$\text{pH} = -\log[H^+(aq)] = -\log(1.643\times10^{-3}) = 2.78$$

The $[H^+(aq)]$ is 1.6×10^{-3} mol/L and the pH is 2.78.

5. **Textbook question 32:**

Use the expression K_a to calculate the $[H^+_{(aq)}]$ because the [acid] is not equal to 0.10 mol/L.

$$C_6H_4COOCH_3COOH_{(aq)} \rightleftarrows C_6H_4COOCH_3COO^-_{(aq)} + H^+_{(aq)}$$

$$K_a = \frac{[H^+_{(aq)}][C_6H_4COOCH_3COO^-_{(aq)}]}{[C_6H_4COOCH_3COOH_{(aq)}]}$$

$$= \frac{[H^+_{(aq)}]^2}{[C_6H_4COOCH_3COOH_{(aq)}]}$$

$$[H^+_{(aq)}] = \sqrt{(3.27\times10^{-4}\ \text{mol/L})(0.018\ \text{mol/L})}$$
$$= 2.426\times10^{-3}\ \text{mol/L}$$
$$= 2.4\times10^{-3}\ \text{mol/L}$$

$$pH = -\log[H^+_{(aq)}]$$
$$= -\log(2.426\times10^{-3})$$
$$= 2.62$$

The $[H^+_{(aq)}]$ is 2.4×10^{-3} mol/L and the pH is 2.62.

The pH would probably decrease (become more acidic) as the temperature increases to 37°C because at a higher temperature more acetylsalicylic acid would dissolve (degree of dissolving of solids generally increases with an increase in temperature), pushing the equilibrium to the right to produce more $H^+_{(aq)}$ ions.

Textbook question 33:

The $[H^+_{(aq)}]$ is needed to calculate the pH. Use the K_a expression to calculate the $[H^+_{(aq)}]$ because the [acid] = 0.50 mol/L.

$$H_3BO_{3(aq)} \rightleftarrows H^+_{(aq)} + H_2BO_3^-{}_{(aq)}$$

$$K_a = 5.8\times10^{-10}\ \text{mol/L}$$
$$= \frac{[H^+_{(aq)}][H_2BO_3^-{}_{(aq)}]}{[H_3BO_{3(aq)}]}$$

$$[H^+_{(aq)}] = \sqrt{(5.8\times10^{-10}\ \text{mol/L})(0.50\ \text{mol/L})}$$
$$= 1.703\times10^{-5}\ \text{mol/L}$$
$$= 1.7\times10^{-5}\ \text{mol/L}$$

$$pH = -\log[H^+_{(aq)}]$$
$$= -\log(1.703\times10^{-5})$$
$$= 4.77$$

The pH of the solution is 4.77.

6. The necessary calculations are as follows:

Ⓐ $[H^+(aq)] = 10^{-pH}$ Ⓑ $K_a = \dfrac{[H^+(aq)]^2}{[acid]}$

7.

Acid	[acid] (mol/L)	pH	$[H^+(aq)]$ (mol/L)	K_a (mol/L)
$HCN_{(aq)}$	0.010	5.60	2.5×10^{-6}	6.3×10^{-10}
$HCOOH_{(aq)}$	0.25	2.17	6.8×10^{-3}	1.8×10^{-4}
$CH_3COOH_{(aq)}$	0.15	2.78	1.7×10^{-3}	1.8×10^{-5}
$HClO_{(aq)}$	0.050	4.42	3.8×10^{-5}	2.9×10^{-8}
$HNO_{2(aq)}$	0.12	2.03	9.3×10^{-3}	7.3×10^{-4}

Sample calculations for $HCN_{(aq)}$ are as follows:

$HCN_{(aq)} \rightleftarrows H^+(aq) + CN^-(aq)$

$[H^+(aq)] = 10^{-pH}$

$= 10^{-5.60}$

$= 2.512\times10^{-6}$ mol/L

$= 2.5\times10^{-6}$ mol/L

$K_a = \dfrac{[H^+(aq)]^2}{[HCN_{(aq)}]}$

$= \dfrac{(2.512\times10^{-6}\ \text{mol/L})^2}{0.010\ \text{mol/L}}$

$= 6.3\times10^{-10}$ mol/L

8. **Textbook question 34:**

$C_6H_4OHCOOH_{(aq)} \rightleftarrows H^+(aq) + C_6H_4OHCOO^-(aq)$

$K_a = \dfrac{[H^+(aq)][C_6H_4OHCOO^-(aq)]}{[C_6H_4OHCOOH^-(aq)]}$

$[C_6H_4OHCOOH_{(aq)}] = \dfrac{1\ \text{g}}{0.460\ \text{L}} \times \dfrac{1\ \text{mol}}{138.13\ \text{g}}$

$= 0.01574$ mol/L

$= 0.02$ mol/L

$[H^+(aq)] = [C_6H_4OHCOO^-(aq)]$

$= 10^{-pH}$

$= 10^{-2.4}$

$= 3.981\times10^{-3}$ mol/L

$= 4\times10^{-3}$ mol/L

$K_a = \dfrac{[H^+(aq)]^2}{[C_6H_4OHCOOH_{(aq)}]}$

$= \dfrac{(3.981\times10^{-3}\ \text{mol/L})^2}{(0.01574\ \text{mol/L})}$

$= 1.007\times10^{-3}$ mol/L

$= 1\times10^{-3}$ mol/L

The K_a value for salicylic acid is 1×10^{-3} mol/L.

9. $[\text{acid}] = \dfrac{\left[H^+_{(aq)}\right]^2}{K_a}$

10. $HCOOH_{(aq)} \rightleftarrows HCOO^-_{(aq)} + H^+_{(aq)}$

$$K_a = \frac{\left[H^+_{(aq)}\right]\left[HCOO^-_{(aq)}\right]}{\left[HCOOH_{(aq)}\right]}$$

$$\begin{aligned}\left[H^+_{(aq)}\right] &= 10^{-pH} \\ &= 10^{-2.33} \\ &= 4.677 \times 10^{-3} \text{ mol/L} \\ &= 4.7 \times 10^{-3} \text{ mol/L}\end{aligned}$$

$$\begin{aligned}\left[HCOOH_{(aq)}\right] &= \frac{\left[H^+_{(aq)}\right]^2}{K_a} \\ &= \frac{\left(4.677 \times 10^{-3} \text{ mol/L}\right)^2}{1.8 \times 10^{-4} \text{ mol/L}} \\ &= 0.1215 \text{ mol/L} \\ &= 0.12 \text{ mol/L}\end{aligned}$$

The $\left[HCOOH_{(aq)}\right]$ is 0.12 mol/L.

11. **Textbook question 28:**

For each of the questions a-c, $\left[H^+_{(aq)}\right] = \dfrac{(\text{percent ionization})}{100}[\text{acid}]$

a. For 0.10 mol/L $HF_{(aq)}$, $\left[H^+_{(aq)}\right] = (7.8/100)(0.10 \text{ mol/L})$
$= 7.8 \times 10^{-3}$ mol/L

b. For 2.3 mmol/L $HNO_{3(aq)}$, $\left[H^+_{(aq)}\right] = (100/100)\left(2.3 \times 10^{-3} \text{ mol/L}\right)$
$= 2.3 \times 10^{-3}$ mol/L (This is an approximation, since the concentration is not 0.10 mol/L.)

c. For 0.10 mol/L $HCN_{(aq)}$, $\left[H^+_{(aq)}\right] = (0.0078/100)(0.10 \text{ mol/L})$
$= 7.8 \times 10^{-6}$ mol/L

d. The term *most acidic* refers to the highest $\left[H^+_{(aq)}\right]$. The 0.10 mol/L $HF_{(aq)}$ is the most acidic (it has the highest $\left[H^+_{(aq)}\right]$). Even though $HF_{(aq)}$ is a weak acid, the strong acid $HNO_{3(aq)}$ has such a low concentration that its solution is less acidic.

Textbook question 29:

$$\% \text{ ionization} = \frac{\left[H^+_{(aq)}\right]}{\left[C_2H_5COOH_{(aq)}\right]} \times 100$$

$$= \frac{\left(1.16 \times 10^{-3} \ \cancel{\text{mol/L}}\right)}{\left(0.100 \ \cancel{\text{mol/L}}\right)} \times 100$$

$$= 1.16\%$$

The percent ionization (reaction) of propanoic acid is 1.16%.

Textbook question 30:

$$\left[H^+_{(aq)}\right] = 10^{-pH}$$

$$= 10^{-2.43}$$

$$= 3.715 \times 10^{-3} \text{ mol/L}$$

$$= 3.7 \times 10^{-3} \text{ mol/L}$$

$$\% \text{ ionization} = \frac{\left[H^+_{(aq)}\right]}{[\text{acid}]} \times 100$$

$$= \frac{3.715 \times 10^{-3} \ \cancel{\text{mol/L}}}{0.10 \ \cancel{\text{mol/L}}} \times 100$$

$$= 3.7\%$$

The percent ionization of lactic acid is 3.7%.

Textbook question 31:

$$HF_{(aq)} \rightleftarrows H^+_{(aq)} + F^-_{(aq)}$$

$$K_a = \frac{\left[H^+_{(aq)}\right]\left[F^-_{(aq)}\right]}{\left[HF_{(aq)}\right]}$$

$$\left[H^+_{(aq)}\right] = \sqrt{\left(6.6 \times 10^{-4} \text{ mol/L}\right)\left(2.0 \text{ mol/L}\right)}$$

$$= 3.633 \times 10^{-2} \text{ mol/L}$$

$$= 3.6 \times 10^{-2} \text{ mol/L}$$

$$\left[H^+_{(aq)}\right] = \left[F^-_{(aq)}\right]$$

$$\left[F^-_{(aq)}\right] = 3.6 \times 10^{-2} \text{ mol/L}$$

The $\left[H^+_{(aq)}\right]$ is 3.6×10^{-2} mol/L; the $\left[F^-_{(aq)}\right]$ is also 3.6×10^{-2} mol/L.

12. a. All ionic hydroxides are considered strong bases because they all dissociate 100% when dissolved in water, even though they dissolve to different degrees.

b. Many hydroxides are low in solubility and do not form solutions with high concentrations.

13. $NaOH_{(aq)} \rightleftarrows Na^{+}_{(aq)} + OH^{-}_{(aq)}$

For 0.10 mol/L $NaOH_{(aq)}$, $[OH^{-}_{(aq)}] = 0.10$ mol/L, the pOH = 1.00, and the pH = 13.00

$$pOH = -\log[OH^{-}_{(aq)}]$$
$$= -\log(0.10)$$
$$= 1.00$$

$$pH = 14.00 - pOH$$
$$= 14.00 - 1.00$$
$$= 13.00$$

$Ba(OH)_{2(aq)} \rightleftarrows Ba^{2+}_{(aq)} + 2\ OH^{-}_{(aq)}$

For 0.10 mol/L $Ba(OH)_{2(aq)}$, $[OH^{-}_{(aq)}] = 2 \times 0.10$ mol/L
$= 0.20$ mol/L

$$pOH = -\log[OH^{-}_{(aq)}]$$
$$= -\log(0.20)$$
$$= 0.6990$$
$$= 0.70$$

$$pH = 14.00 - pOH$$
$$= 14.00 - 0.6990$$
$$= 13.30$$

14. Based on the evidence, Solutions 1 and 6 are basic and must be $Ba(OH)_{2(aq)}$ and $KOH_{(aq)}$ since they turn litmus blue. Solution 1 has a higher conductivity and must therefore be $Ba(OH)_{2(aq)}$. Solution 6 must then be $KOH_{(aq)}$. Solutions 3 and 5 must be neutral because they have no effect on red or blue litmus. Since Solution 3 does not conduct electricity, it must be $C_{12}H_{22}O_{11(aq)}$. Solution 5 must be $NaCl_{(aq)}$, the other neutral solution. Solutions 2 and 4 are acidic because they turn blue litmus red. Since Solution 4 conducts much better, it must be the strong acid $HBr_{(aq)}$, whereas Solution 2 is the weak acid $CH_3COOH_{(aq)}$.

15. **Textbook question 26:**

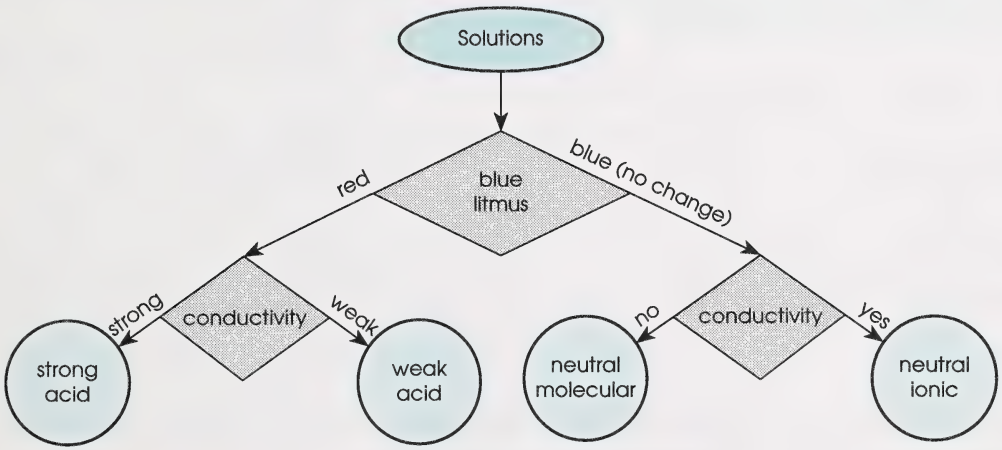
Other empirical properties could have been used to determine the identities of the unknown substances. Each solution could have been tested with an active metal such as magnesium; the two acids would have reacted and the faster, more vigorous reaction would have identified the stronger acid. The other four solutions could be tested with $Mg(NO_3)_{2(aq)}$; the bases would have formed insoluble $Mg(OH)_{2(s)}$. If equal volumes of solutions were tested, the $Ba(OH)_{2(aq)}$ would form more precipitate since there are more $OH^{-}_{(aq)}$ ions. The two remaining solutions could be tested with $AgNO_{3(aq)}$; the $NaCl_{(aq)}$ forms a precipitate, $AgCl_{(s)}$.

Textbook question 30:

The experimental design may be as follows:

Samples of four solutions are tested to classify as strong acid, weak acid, neutral ionic, or neutral molecular. Each solution is tested first with blue litmus paper; then for conductivity. The temperatures and concentrations of solutions are kept constant.

A flow chart of the experimental design may look as follows:



Section 2: Follow-up Activities

Extra Help

1. a. x c. viii e. iii g. vii i. iv
   b. v d. ii f. i h. ix j. vi

2. The explanations are as follows:

   a. Determine pOH using pH.

   b. Convert pOH into $\left[OH^-_{(aq)}\right]$.

   c. Find the moles of $OH^-_{(aq)}$ by multiplying the $\left[OH^-_{(aq)}\right]$ by the given volume.

   d. Find the moles of $NaOH_{(s)}$ by using the mole to mole ratio from the dissociation equation.

   e. Find the mass of $NaOH_{(s)}$ by multiplying the moles of $NaOH_{(s)}$ by its molar mass.

3. If the pH is <7, then the pOH is >7 and the substance is an **acid**. If the pOH is <7, then the pH is >7 and the substance is a **base**.

4.

Diagnostic Test Results	Substance
turned litmus red, conducted well	$HCl_{(aq)}$ (strong acid)
turned litmus blue, conducted well	$Ba(OH)_{2(aq)}$ (strong base)
no change in litmus, conducted well	$NaCl_{(aq)}$ (neutral ionic)
turned litmus red, poor conductor	$CH_3COOH_{(aq)}$ (weak acid)
no change in litmus, no conductivity	$C_6H_{12}O_{6(aq)}$ (neutral molecular)

5. The two methods of calculating $\left[H^+_{(aq)}\right]$ are as follows:

$$\left[H^+_{(aq)}\right] = \sqrt{K_a[\text{acid}]} \qquad \textbf{and} \qquad \left[H^+_{(aq)}\right] = \frac{\%\text{ ionization}}{100} \times [\text{acid}]$$

The K_a method works when the [acid] is at least 1000 times greater than the K_a value.

The percent ionization method works if the [acid] is 0.10 mol/L. The method also works for strong acids at other concentrations.

Enrichment

1. The approximation of K_a has to be adjusted. The [acid] must be decreased by the amount of acid that ionized, which equals $\left[H^+_{(aq)}\right]$. The K_a expression then becomes

$$K_a = \frac{\left[H^+_{(aq)}\right]^2}{\left([\text{acid}] - \left[H^+_{(aq)}\right]\right)} \qquad \textbf{or} \qquad K_a = \frac{x^2}{([\text{acid}] - x)}, \text{ where } x = \left[H^+_{(aq)}\right]$$

To solve for the $\left[H^+_{(aq)}\right]$, and then the pH of a 0.20 mol/L $H_2SO_{3(aq)}$, the quadratic equation must be used.

$$1.3 \times 10^{-2} = \frac{x^2}{0.20 - x}$$

$$x^2 = 2.6 \times 10^{-3} - 1.3 \times 10^{-2} x$$

$$x^2 + 1.3 \times 10^{-2} x - 2.6 \times 10^{-3} = 0$$

Therefore, $a = 1$

$b = 1.3 \times 10^{-2}$

$c = -2.6 \times 10^{-3}$

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

$$x = \frac{-1.3 \times 10^{-2} \pm \sqrt{\left(1.3 \times 10^{-2}\right)^2 - 4(1)\left(-2.6 \times 10^{-3}\right)}}{2}$$

$$= \frac{-1.3 \times 10^{-2} \pm \sqrt{1.057 \times 10^{-2}}}{2}$$

$$= \frac{-1.3 \times 10^{-2} \pm 1.028 \times 10^{-1}}{2}$$

$$= 4.49 \times 10^{-2} \quad \textbf{or} \quad -5.79 \times 10^{-2}$$

$$\left[H^+_{(aq)}\right] = \boxed{4.5 \times 10^{-2} \text{ mol/L}} \textbf{ or } \boxed{-5.8 \times 10^{-2} \text{ mol/L}}$$

You cannot have a negative concentration, so the only possible answer for the $\left[H^+_{(aq)}\right]$ is 4.5×10^{-2} mol/L.

$$pH = -\log\left[H^+_{(aq)}\right]$$
$$= -\log\left(4.49 \times 10^{-2}\right)$$
$$= 1.35$$

The pH of the solution is 1.35.

2. Reaction 1: $CO_{2(aq)} + H_2O_{(l)} \rightleftarrows H^+_{(aq)} + HCO_3^-{}_{(aq)}$

Reaction 2: $H^+_{(aq)} + CaCO_{3(s)} \rightleftarrows Ca^{2+}_{(aq)} + HCO_3^-{}_{(aq)}$

a. These two reactions are tied together very closely. Any water flowing on the earth picks up $CO_{2(g)}$ from the air, dissolves it, and then reacts with it, as shown in Reaction 1. The higher the $\left[CO_{2(aq)}\right]$ (dissolved $CO_{2(g)}$), the farther the equilibrium of Reaction 1 is pushed to the right. This creates a greater $\left[H^+_{(aq)}\right]$ in the water. When this water containing increased $\left[H^+_{(aq)}\right]$ flows over limestone $\left(CaCO_{3(s)}\right)$ deposits, as in many riverbeds, it forces the equilibrium of Reaction 2 to the right. This dissolves the $CaCO_{3(s)}$.

When the water containing $CaCO_{3(aq)}$ flows into a cave and drips from the ceiling, the water is exposed to air, which evaporates some of the water. This increases the concentration of $CaCO_{3(aq)}$. The low solubility of calcium carbonate then precipitates the $CaCO_{3(s)}$ into the formations that result.

b. Acidified rain acts on the Reaction 2 equilibrium. The increased $\left[H^+_{(aq)}\right]$ pushes the equilibrium to the right to use up the excess $\left[H^+_{(aq)}\right]$ and at the same time reacts with the $CaCO_{3(s)}$ (limestone) in the buildings and statues. This equilibrium has had a devastating effect on areas such as the European continent where there is a high concentration of industry. The industrial emissions create acid rain through equilibria such as the one shown by Reaction 1. The limestone structures then deteriorate rapidly through the Reaction 2 equilibrium.

c. Having chickens drink water containing $CO_{2(g)}$ increases the $CO_{2(aq)}$ in the blood, forcing the Reaction 1 equilibrium to the right. Body buffers neutralize the increased $\left[H^+_{(aq)}\right]$ which leaves an increased $\left[HCO_3^-{}_{(aq)}\right]$. This higher $\left[HCO_3^-{}_{(aq)}\right]$ forces the Reaction 2 equilibrium to the left, which precipitates more $CaCO_{3(s)}$ into the eggshell structure, and makes the shells thicker and stronger.

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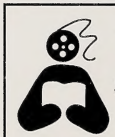

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Module 5

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